Name: ______________________________ Section: ____________________

Workshop #1: Measurements & Conversions

1. Round the following numbers to THREE significant figures, and express your final responses using scientific notation.

   A. 239,720 ____________________ C. 0.000238505 ____________________

   B. 0.09763400 ____________________ D. 7,689,994,656 ____________________

2. Round the following numbers to FOUR significant figures, and express your final responses using scientific notation.

   A. 0.00765796 ____________________ C. 423.56 ____________________

   B. 56,928.31 ____________________ D. 0.0000555226 ____________________

3. Solve each of the following problems. Express your final answer to the correct number of significant figures in scientific notation. Make certain to include the appropriate units where appropriate.

   A. 382.5 mL + 96.31 mL – 5.9 mL ____________________

   B. $\dfrac{3.496\text{ ft} + 27.22\text{ ft}}{5.006\text{ lb}}$ ____________________

   C. $\dfrac{(2.661 \times 10^{-3}\text{ cm})(5.11 \times 10^{9}\text{ cm})}{7.3 \times 10^{7}\text{ cm}}$ ____________________

   D. $\dfrac{28.62\text{ s} - 3.5\text{ s}}{(32.9 \times 10^{2}\text{ s})(99.55 \times 10^{6}\text{ s})}$ ____________________

   E. $\dfrac{(6.345 \times 10^{-17})(2.6447 \times 10^{-45})}{4.567 \times 10^{5} + 7.89887 \times 10^{6}}$ ____________________

4. Solve the following problems, conforming to the appropriate number of significant figures. You may need your textbook for certain unit conversions:

   A. ____________ How many centimeters are there in 3.0 miles?

   B. ____________ Convert $9.06 \times 10^{6}\ \mu m^{2}$ to mm^{2}.

Name: ________________________________ Section: ____________________

C. ____________ Convert 45 meters per second to kilometers per hour.

D. ____________ Determine the density (in g/mL) of a substance that weighs 0.695 lb and occupies a volume of 3.4 qt.

E. ____________ The concentration of carbon monoxide (CO), a common air pollutant, is found in a room to be 5.7×10^{-3} mg / cm^3. How many grams of CO are present in the room if the room’s dimensions measure 3.5 m × 3.0 m × 3.2 m?

F. ____________ A cylindrical piece of metal is 2.03 inches high, has a diameter of 17.0 mm wide, and weighs 31.599 g. Determine its density. Will this object sink or float in water? Volume (cylinder) = $\pi r^2 h$

G. ____________ Zinc sulfide is treated with sulfuric acid, resulting in a solution with some undissolved bits of zinc sulfide and releasing hydrogen sulfide gas. If 10.85 g of zinc sulfide is treated with 50.00 mL of sulfuric acid (density = 1.153 g / mL), 65.15 g of solution plus undissolved solid remain. What is the volume (in L) of hydrogen sulfide gas evolved from this reaction? The density of hydrogen sulfide gas is 1.393 g / L.

Name: ________________________________ Section: ____________________

5. *APPLICATION!* Nanotechnology, the field of building microscale structures one atom at a time, has progressed in recent years. One potential application of nanotechnology is the construction of artificial cells. The simplest cells could mimic red blood cells, the body's oxygen transporters. For example, nanocontainers, perhaps constructed of carbon, could be pumped full of oxygen and injected into a person's bloodstream. If the person needed additional oxygen, these containers could slowly release oxygen into the blood, allowing tissues that would otherwise die to remain alive. Suppose that nanocontainers were cubic and had an edge length of 25 nanometers.

   A. ____________What is the volume (in L) of one nanocontainer?

   B. ____________ Suppose that each nanocontainer could contain pure oxygen pressurized to a density of 85 g / L. How many grams of oxygen could be contained by each nanocontainer?

   C. ____________ Normal air contains about 0.28 g of oxygen per liter. An average human inhales about 0.50 L of air per breath and takes about 20 breaths per minute. How many grams of oxygen does a human inhale per hour?

   D. ____________What is the minimum number of nanocontainers that a person would need in their bloodstream to provide 1.0 hour's worth of oxygen?

Name: ______________________ Section: ______________

Workshop #2: Safety Data Sheet

Chemicals and other hazardous materials are an integral component of the laboratory environment. A Safety Data Sheet (SDS) provides both workers and emergency personnel with the proper procedures for handling a particular substance. SDS's include information such as physical data, toxicity, health effects, first aid, reactivity, storage, disposal, protective equipment, and spill/leak procedures. SDS's vary in length from 1 to 10 pages, with most being 2 to 4 pages.

SDS's are not intended for use by the general consumer that occasionally works with a substance. Rather, SDS's are for employees who may be occupationally exposed to a hazard at work (40 hrs/week or confined spaces), employers who need to know proper storage and handling, and emergency responders.

In the U.S., the Occupational Safety and Health Administration (OSHA) requires that SDS's be available to employees for potentially harmful substances handled in the workplace under the Hazard Communication regulation. OSHA defines a hazardous chemical as any liquid, solid, or gas that could present a physical or health hazard to an employee.

OSHA requires each department such as chemistry, biology, photography, and ceramics to maintain Safety Data Sheets readily available for employee viewing. There are numerous websites that offer SDS's; some are free while most charge. When chemicals are shipped, they are accompanied by a SDS. Also note that the National Fire Protection Association (NFPA) ratings are the blue, red, yellow, and white diamond labels you see on many hazardous chemical containers.

HEALTH HAZARD-BLUE:
- 4 – Deadly
- 3 – Extreme Danger
- 2 – Hazardous
- 1 – Slightly Hazardous
- 0 – Normal Material

REACTIVITY-YELLOW:
- 4 – May Detonate
- 3 – Shock and Heat May Detonate
- 2 – Violent Chemical Change
- 1 – Unstable if Heated
- 0 – Stable

FIRE HAZARD (flash points)-RED:
- 4 – Below 73°F
- 3 – Below 100°F
- 2 – Above 100°F not Exceeding 200°F
- 1 – Above 200°F
- 0 – Will Not Burn

SPECIFIC HAZARD-WHITE:
- OX – Oxidizer
- ACID – Acid
- ALK – Alkali (Base)
- COR – Corrosive
- ₩ – Use NO Water

Name: ______________________________ Section: ____________________

Use the SDS provided in lab to answer the following questions:

1. List other names that are synonyms of sodium hydroxide and its formula.

2. What is its melting point?

3. What is done in case of contact with eyes?

4. How should a small spill be handled?

5. What procedure should be done if the substance is swallowed?

6. What are the NFPA Ratings for Health? Fire? Reactivity? Specific Hazard?

7. List three chemicals that should not be stored with NaOH.

8. How should solid NaOH be properly stored?

Name: ______________________________ Section: ____________________

Your instructor will assign you a specific chemical compound along with SDS, and you should fill-in the table below with as much information as possible (note: several areas will remain blank) using the various resources listed below:

Name of Substance __________________ Chemical Formula ______________

	Reagent Bottle	SDS Sheet	Merck Index	CRC or Lange's
Other Names				
Formula Weight				
State of matter				
Melting point				
Boiling point				
Density				
Percent Composition				
Soluble solvents				
Manufacturer				
Chemical Properties				
Toxicity				

Contrast the differences between the four reference materials used above, and be specific.

Name: ___________________________________ Section: ______________________

Workshop #3: Nomenclature

A. Provide a chemical name for the following formulas:

1. $CuSO_3$ ____________________
2. Hg_2Cl_2 ____________________
3. $BaCr_2O_7$ ____________________
4. NO ____________________
5. $Sr(OH)_2$ ____________________
6. $Mn(NO_2)_2$ ____________________
7. $NaHCO_3$ ____________________
8. $HNO_3(aq)$ ____________________
9. $CsClO_2$ ____________________
10. Ag_3PO_3 ____________________
11. $V_2(CrO_4)_5$ ____________________
12. $Sn(MnO_4)_4$ ____________________
13. I_2O_7 ____________________

Name: ______________________________ Section: ____________________

B. Provide a formula for the following names:

1. sodium peroxide ______________________________

2. copper(II) sulfate pentahydrate ______________________________

3. ammonia ______________________________

4. sulfurous acid ______________________________

5. calcium hydride ______________________________

6. ammonium hydrogen phosphate ______________________________

7. arsenic(III) sulfate ______________________________

8. dichlorine heptoxide ______________________________

9. gold(I) iodide ______________________________

10. antimony(III) nitride ______________________________

11. tin(IV) carbonate ______________________________

12. bismuth(III) oxide ______________________________

13. mercury(II) perchlorate ______________________________

14. pentane ______________________________

Name: ___________________________________ Section: ______________________

C. Provide a chemical name for the following formulas:

1. $HC_2H_3O_2(aq)$ ______________________________

2. PbC_2O_4 ______________________________

3. $Au(ClO)_3$ ______________________________

4. $Cd(SCN)_2$ ______________________________

5. $CuMnO_4$ ______________________________

6. KIO_3 ______________________________

7. ClO_2 ______________________________

8. TiH_4 ______________________________

9. HCl(g) ______________________________

10. $As(HSO_4)_3$ ______________________________

11. SO_3 ______________________________

12. $Fe(OH)_2$ ______________________________

13. $CH_3CH_2CH_2CH_3$ ______________________________
(i.e. C_4H_{10})

14. $CH_3(CH_2)_6CH_3$ ______________________________
(i.e. C_8H_{18})

Name: ______________________________ Section: ___________________

D. Provide a formula for the following names:

1. tungsten(V) phosphide ______________________________
2. gallium nitrate ______________________________
3. carbonic acid ______________________________
4. xenon hexachloride ______________________________
5. hydrosulfuric acid ______________________________
6. lithium dihydrogen phosphite ______________________________
7. nonane ______________________________
8. lead(IV) oxalate ______________________________
9. phosphoric acid ______________________________
10. dinitrogen tetroxide ______________________________
11. sodium selenate ______________________________
12. sodium bicarbonate ______________________________
13. hypobromous acid ______________________________
14. zinc oxide ______________________________

Name: ______________________________ Section: ____________________

Workshop #4: Reactions

Predict products and balance the following reactions (write total-ionic and net-ionic where requested). If no reaction takes place, write NR for no reaction. Be sure to include phases.

1. Synthesis (Combination or Composition) Reactions: $A + B \rightarrow AB$

   A. $Ca(s) + N_2(g) \rightarrow$

   B. $SrO(s) + H_2O(l) \rightarrow$

   C. $N_2O_5(s) + H_2O(l) \rightarrow$

   D. $CaO(s) + CO_2(g) \rightarrow$

2. Decomposition Reactions: $AB \rightarrow A + B$

   A. $HgO(s) \rightarrow$

   B. $MgSO_4{\cdot}7H_2O(s) \rightarrow$

   C. $KClO_3(s) \rightarrow$

   D. $BaCO_3(s) \rightarrow$

3. Combustion Reactions: *nonmetals + O_2 → nonmetal oxides: H_2O, CO_2, SO_2, NO_2*

   A. $CH_3CH_2OH(l) + O_2(g) \rightarrow$

   B. $B_2H_6(g) + O_2(g) \rightarrow$

   C. $NH_3(g) + O_2(g) \rightarrow$

   D. $C_{10}H_{22}(g) + O_2(g) \rightarrow$

Name: ________________________________ Section: ____________________

4. Single Replacement (Displacement) Reactions: *C* + *AB* → *AC* + *B* <u>*OR*</u> *CB* + *A*

A. *molecular*: $Al(s)$ + $CuCl_2(aq)$ →

total-ionic:

net-ionic:

B. *molecular*: $H_2(g)$ + $Fe_2O_3(s)$ →

total-ionic:

net-ionic:

C. *molecular*: $Cl_2(g)$ + $NaBr(aq)$ →

total-ionic:

net-ionic:

D. *molecular*: $Na(s)$ + $H_2O(l)$ →

total-ionic:

net-ionic:

5. Double Replacement (Displacement) Reactions: *AB* + *CD* → *AD* + *CB*

A. *molecular*: $Ba(OH)_2(aq)$ + $Fe_2(SO_4)_3(aq)$ →

total-ionic:

net-ionic:

B. *molecular*: $HCl(aq)$ + $Hg_2(NO_3)_2(aq)$ →

total-ionic:

net-ionic:

C. *molecular*: $CaCO_3(s)$ + $HC_2H_3O_2(aq)$ →

total-ionic:

net-ionic:

D. *molecular*: $HClO_4(aq)$ + $Fe(OH)_3(s)$ →

total-ionic:

net-ionic:

6. Redox (Oxidation-Reduction) Reactions:

A. $As_2O_3(s)$ + $NO_3^-(aq)$ → $AsO_4^{-3}(aq)$ + $NO(g)$ (under acidic conditions)

Oxidation half reaction:

Reduction half reaction:

Balanced reaction:

B. $Cl_2(g)$ → $Cl^-(aq)$ + $ClO^-(aq)$ (under basic conditions)

Oxidation half reaction:

Reduction half reaction:

Balanced reaction:

Name: ______________________________ Section: ____________________

Workshop #5: Stoichiometry

Show calculation setups and answers for all problems below.

1. How many molecules are there in a 600.0 g sample of Na_3PO_4(s)? How many Na^+ ions are present?

2. A compound of copper and sulfur was produced in the lab by heating copper and sulfur together in a crucible. The following data was collected:

Mass of crucible and cover	28.71 g
Mass of crucible, cover, and copper	30.25 g
Mass of crucible, cover, and copper-sulfur compound	30.64 g

   Determine the empirical formula of this compound.

3. Isopentyl acetate ($C_7H_{14}O_2$), the compound responsible for the scent of bananas, can be produced commercially. Calculate the percent composition of $C_7H_{14}O_2$.

Name: ______________________________ Section: ____________________

4. A compound consisting of mainly cetyl palmitate is comprised entirely of carbon, hydrogen, and oxygen. Combustion of a 2.3836 g sample of cetyl palmitate produced 6.9807 g of CO_2 and 2.8575 g of H_2O. Determine the empirical formula of the compound. If the formula weight of the compound is 480.9 g/mol, what is the molecular formula of this compound?

5. Washing soda, a compound used to prepare hard water for laundry, is a hydrate whose formula can be written as $Na_2CO_3 \cdot xH_2O$. When a 2.558 g sample of washing soda is heated at 125 °C, all the water of hydration is lost, leaving behind 0.948 g of the anhydrous salt. Determine the value of x.

Name: ________________________________ Section: ____________________

6. Liquid mercury and bromine gas will react under appropriate conditions to produce solid mercury(II) bromide.

   A. Write the balanced chemical equation for this process.
   B. What is the maximum mass of $HgBr_2$ that can be produced from the reaction of 10.0 g Hg and 9.00 g Br_2?
   C. Determine the remaining mass of each reactant (if any) available upon conclusion of the reaction.
   D. If 15.3 g of mercury(II) bromide is produced in this reaction, determine the percentage yield of product.

7. Silicon nitride (Si_3N_4), a valuable ceramic, is made by the direct combination of silicon and nitrogen at high temperature.

   A. Write the balanced chemical equation for this process.
   B. How many grams of silicon must react with excess nitrogen to prepare 125 g silicon nitride if the yield of the reaction is 85.0%?

Name: ________________________________ Section: ____________________

8. Consider the following unbalanced reaction:

$$XNO_3(aq) + CaCl_2(aq) \rightarrow XCl(s) + Ca(NO_3)_2(aq)$$

If 30.8 g of $CaCl_2$ produced 79.6 g of XCl, determine the identity of X. Quantify your response. Random guessing will not earn any credit for this problem!

Name: ______________________________ Section: ____________________

Workshop #6: Solution Stoichiometry

Write balanced equations and show calculation setups for all the problems below.

1. A 1.192 g sample of oxalic acid, $H_2C_2O_4$, is placed in a 100.0 mL volumetric flask and filled to the mark with water. What is the molarity of the solution?

2. How many grams of sodium dichromate should be added to a 50.0 mL volumetric flask to prepare a 0.025 M sodium dichromate solution when the flask is filled to the mark with water?

3. A chemist wants to prepare 0.250 M HCl(aq). Commercial hydrochloric acid is 12.4 M. How many milliliters of the commercial acid does the chemist require to make up 1.50 L of the dilute acid?

4. If 35.4 g of aluminum are treated with 721 mL of 5.86 M HCl, how many grams of hydrogen gas will theoretically be formed?

Name: ________________________________ Section: ____________________

5. The concentration of hydrogen peroxide in a solution is determined by titrating a 10.0 mL sample of the solution with permanganate ion under acidic conditions, producing manganese(II) ion and oxygen gas. If it takes 13.5 mL of 0.109 M MnO_4^- solution to reach the equivalence point, what is the molarity of the hydrogen peroxide solution?

6. A flask contains 49.8 mL of 0.150 M calcium hydroxide solution. How many milliliters of 0.350 M sodium carbonate are required to react completely with the calcium hydroxide?

7. During the developing process of black and white film, silver bromide is removed from photographic film by the fixer. The major component of the fixer is sodium thiosulfate. What mass of silver bromide can be dissolved by 1.00 L of 0.200 M sodium thiosulfate?

$$AgBr(s) + S_2O_3^{-2}(aq) \rightarrow Ag(S_2O_3)_2^{-3}(aq) + Br^-(aq)$$ *(unbalanced)*

Name: ______________________________ Section: ____________________

8. A 3.33 gram sample of iron ore is transformed to a solution of iron(II) sulfate, and this solution is titrated with 0.150 M potassium dichromate. If it required 41.4 mL of potassium dichromate solution to titrate the iron(II) sulfate solution, what is the percentage of iron in the ore?

$$Fe^{+2}(aq) + Cr_2O_7^{-2}(aq) \rightarrow Fe^{+3}(aq) + Cr^{+3}(aq) \qquad \textit{(unbalanced)}$$

Name: ______________________________ Section: ____________________

Workshop #7: Gas Laws

Show calculation setups and answers for all problems below.

1. A particular balloon is designed by its manufacturer to be inflated to a volume of no more than 2.5 L. The balloon is filled with 2.0 L of helium at sea level (pressure = 1.00 atm), is released, and rises to an altitude at which the atmospheric pressure is only 500.0 mmHg. Assuming that the temperature remains constant, will the balloon burst? Quantify your response and briefly explain.

2. Another balloon is filled with 150 L of helium at 23 °C and 1.0 atm. What volume does the balloon have when it has risen to a point in the atmosphere where the pressure is 220 mmHg and the temperature is –31 °C?

3. Calculate the mass of hydrogen gas needed to fill an 80.0 L tank to a pressure of 2205 psi at 27 °C.

4. What volume does 35 mol of nitrogen gas occupy at STP?

5. The mass of a 3.21 L gas is found to be 3.50 g, measured at 65.0 ºC and 500.0 torr. Determine the molar mass of the gas.

Name: ______________________________ Section: ____________________

6. Calculate the density of water vapor at 110 °C and 99 kPa.

7. A compound with the empirical formula BH_3 was found to have a vapor density of 1.24 g / L at STP. Determine the molecular weight AND the molecular formula of this gas.

8. Consider the reaction of solid copper(I) sulfide with oxygen gas to produce solid copper(I) oxide and gaseous sulfur dioxide.

   A. Write the balanced chemical equation for this process.

   B. What volume of oxygen gas, measured at 27.5 °C and 0.998 atm, is required to react with 25 g of copper(I) sulfide?

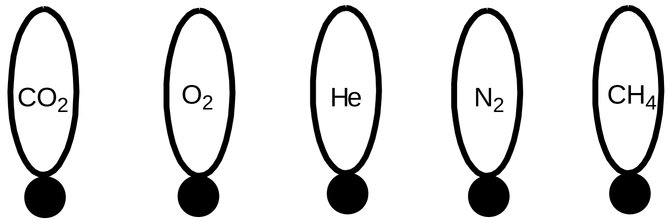
9. A sample of solid potassium chlorate is decomposed, forming solid potassium chloride and gaseous oxygen. The oxygen produced was collected by displacement of water at 22 °C at a total pressure of 754 torr. The volume of the gas collected was 0.65 L, and the vapor pressure of water at 22 °C is 21 torr.

   A. Write the balanced chemical equation for this process.

   B. Determine the mass of potassium chlorate in the sample that was decomposed.

Name: ________________________________ Section: ____________________

10. Represented below are five identical balloons, each filled to the same volume at 25 °C and 1.0 atm pressure with the pure gases indicated.



A. Which balloon contains the greatest mass of gas? Explain.

B. Compare the average kinetic energies of the gas molecules in the balloons. Explain.

C. Which balloon contains the gas that would be expected to deviate most from the behavior of an ideal gas? Explain.

D. Twelve hours after being filled, all the balloons have decreased in size. Predict which balloon will be the smallest. Explain your reasoning.

11. Calculate the root-mean-square speed (u_{rms}) for:

A. a xenon atom at 298 K;

B. an oxygen molecule at 298 K.

12. Both hydrogen and helium have been used as buoyant gases in blimps. If a small leak were to occur in a blimp filled with both gases, which gas would effuse more rapidly and by what factor?

Name: ________________________________ Section: ____________________

13. A gas of unknown molecular mass was allowed to effuse through a small opening under constant pressure conditions. It required 72 s for the gas to effuse. Under identical experimental conditions, it required 28 s for O_2 gas to effuse. Determine the molar mass of the unknown gas.

14. Calculate the pressure exerted by 50.0 g CO(g) in a 1.00 L container at 25 °C by:

    Useful information: For CO, a = 1.49 atm L^2/mol^2 and b = 0.0399 L/mol

    A. using the ideal gas law, and

    B. using the van der Waals equation.

15. Compare the results from parts A and B. Does CO(g) behave ideally under these conditions? Briefly explain why or why not.

Name: ______________________________ Section: ____________________

Workshop #8: Thermochemistry

Show calculation setups and answers for each question. *Please note that your instructor may opt to assign specific questions from those listed below.*

1. Calculate the change in internal energy (in J) for a balloon that is heated by adding 215 cal of heat. It expands, doing 422 J of work on the atmosphere.

2. Consider the following *balanced* reaction: $CH_3OH(g) \rightarrow CO(g) + 2H_2(g)$, where $\Delta H = +90.7$ kJ. If the enthalpy change is 16.5 kJ, how many grams of hydrogen gas are produced?

3. A 50.00 g sample of an unknown substance absorbed 2.578 kJ of energy as it changed from a temperature of 25.0 ºC to 89.7 ºC. What is the specific heat of this unknown substance (in J/g ºC)?

4. An alloy of mass 25.0 g was heated to 88.6 °C and then placed in a calorimeter that contained 61.2 g of water at 19.6 °C. The temperature of the water rose to 21.3 °C. Determine the specific heat of the alloy (in J/g ºC).

5. 100.0 g of copper metal, initially at 100.0 ºC, is added to a calorimeter containing 250.0 g of H_2O at 15.0 ºC. If the specific heat of copper is 0.389 J/g ºC, what is the final temperature of the water and copper mixture?

Name: ________________________________ Section: ____________________

6. The chemical equation for the combustion of magnesium in sulfur dioxide is

$$3\ Mg(s) + SO_2(g) \rightarrow MgS(s) + 2\ MgO(s)$$

Calculate the $\Delta H°_{rxn}$ (in kJ) given the following thermodynamic data:

$$Mg(s) + \frac{1}{2}\ O_2(g) \rightarrow MgO(s) \qquad \Delta H° = –601.7\ kJ$$
$$Mg(s) + S(s) \rightarrow MgS(s) \qquad \Delta H° = –598.0\ kJ$$
$$S(s) + O_2(g) \rightarrow SO_2(g) \qquad \Delta H° = –296.8\ kJ$$

7. Consider the following thermochemical equation:

$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$$

Determine the $\Delta H°_{rxn}$ (in kcal) given the following thermochemical data:

$$N_2(g) + O_2(g) \rightarrow 2\ NO(g) \qquad \Delta H° = 43.20\ kcal$$
$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g) \qquad \Delta H° = –22.10\ kcal$$
$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g) \qquad \Delta H° = –115.60\ kcal$$

Name: ______________________________ Section: ____________________

8. Consider the neutralization reaction of sodium hydroxide and sulfuric acid in a coffee-cup calorimeter.

$$2\ NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

100.0 mL of 1.00 M aqueous NaOH is mixed with 100.0 mL of 1.00 M aqueous H_2SO_4, each at 24.0 °C, were mixed. The maximum temperature achieved was 30.6 °C. Calculate the enthalpy change of reaction (in kJ/mol) of Na_2SO_4 produced. The specific heat of the reaction is known to be 4.184 J/g ºC. The density of the reaction mixture is 1.00 g/mL. Assume the volumes are additive.

9. Suppose 50.0 mL of HCl is combined with 100.0 mL of 1.05 M NaOH in a coffee-cup calorimeter. The reaction mixture, initially at 22.0 ºC, reached a final temperature of 30.2 ºC. Determine the molarity of the HCl solution assuming all of the HCl reacted and that NaOH is present in excess. The specific heat of the reaction is known to be 0.96 cal/g ºC, and the heat of neutralization is 13.6 kcal/mol. The density of the reaction mixture is 1.02 g/mL. Assume the volumes are additive.

Name: ________________________________ Section: ____________________

Workshop #9: The Atomic Spectrum of Hydrogen

Emission and absorption spectra of hydrogen gave experimental evidence of quantized energy states for electrons within an atom. Niels Bohr tried to explain the observations with the first quantum theory in 1913.

Bohr's Quantum Theory: explains the emission and absorption spectra

1. An atom has discrete energy levels (orbits), where an electron may exist without emitting or absorbing electromagnetic radiation.
2. An electron may move from one orbit to another. By doing so, the electromagnetic radiation is absorbed or emitted.
3. An electron moves in circular orbits about the nucleus, and the energy of the electron is quantized.

Limitation: Bohr's theory works only for an atom or ion containing one electron.

According to this quantum theory, electrons within atoms can only exist in certain states, each of which has an associated fixed amount of energy. When an electron changes its state, it must absorb or emit an amount of energy equal to the difference between the energy of the initial and final states. This energy may be absorbed or emitted in the form of electromagnetic energy (i.e. infrared, visible, or ultraviolet light), in which case the relationship between the change in energy and the wavelength of the light which is associated with the transition is given by the equation:

$$|\Delta E| = \frac{hc}{\lambda} \qquad (1)$$

where $|\Delta E|$ is the absolute value of the change in energy in joules, h is Planck's constant, 6.6262×10^{-34} J·sec, c is the speed of light, 3.00×10^{8} m / sec, and λ is the wavelength. The change in energy, ΔE, of the atom is positive if light is absorbed and negative if it is emitted.

Atomic and molecular spectra are the result of changes in energy which occur in atoms and molecules when they are excited. The emission spectrum of an atom gives us information about the spacing between the allowed energy levels of that atom. The different wavelengths present in the light can be used to establish the energy levels available to the atom. Conversely, given the set of energy levels for an atom, one can predict its atomic spectrum and determine which levels were involved in any observed line in the spectrum.

Ordinarily, an electron will occupy its lowest energy state, called the ground state. If the atom is excited, the electron which is now in a higher energy state will be unstable and will quickly make a transition back down to its ground state. In making the transition, the energy of the atom may be radiated as light.

Name: ______________________________ Section: ____________________

The simplest atomic spectrum is that of the *hydrogen atom*. In 1886, Balmer showed that the lines in the spectrum of the hydrogen atom had wavelengths that could be expressed by a rather simple equation. In 1913, Bohr explained the spectrum on a theoretical basis with his famous model of the hydrogen atom. According to Bohr's theory, the energies E_n allowed to a hydrogen atom are all given by the following equation:

$$E_n = \frac{-B}{n^2} \qquad (2)$$

where B is the constant, 2.178×10^{-18} J and n is an integer, 1, 2, 3,…, called a quantum number. It has been found that all the lines in the atomic spectrum of hydrogen can be associated with differences between atomic energy levels which are predicted with great accuracy by Bohr's equation.

There are several ways in which one might analyze an atomic spectrum, given the energy levels of the atom, but a simple and powerful method is to calculate the wavelengths of some of the lines that are allowed and see if they match those which are observed. We shall use this method in our workshop.

A. Energy Levels of Hydrogen

Given the expression for E_n in Equation 2, calculate the energy (in joules) for each of the levels of the H atom missing in the table below. Notice that the energies are all negative, so that the lowest energy will have the largest allowed negative value. Enter these values in the table of energy levels below:

Table One

Quantum Number	Energy, E_n, in joules	Quantum Number	Energy, E_n, in joules
1	-2.178×10^{-18} J	6	
2		7	
3		8	
4		10	-2.178×10^{-20} J
5		∞	ZERO Joules

(Workshop continued on next page)

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B. Calculation of Wavelengths in the Spectrum of the Hydrogen Atom

The lines in the hydrogen spectrum all arise from jumps made by the atom from one energy level to another. The wavelengths in nanometers of these lines can be calculated by Equation 1, where $|\Delta E|$ is the positive difference in energy between any two allowed levels. By rearranging Equation 1 it is possible to solve for wavelengths:

$$\lambda = \frac{hc}{|\Delta E|} \qquad (3)$$

After putting in some constants we can solve for wavelength in nanometers.

$$\lambda = \frac{(6.626 \times 10^{-34}\,\text{J}\cdot\text{sec})(3.00\times10^{8}\ \text{m / sec})}{|\Delta E|} \times \frac{1\ \text{nm}}{10^{-9}\ \text{m}} \qquad (4)$$

Calculate the $|\Delta E|$ and wavelength for all the jumps indicated in the table below. Write $|\Delta E|$, the difference in energy in J between $E_{n,hi}$ and $E_{n,lo}$, in the upper half of the box, and in the lower half of the box, write the λ (in nm) associated with that value. The box for the $n_2 \rightarrow n_1$ transition is filled in for you.

Table Two

n_{higher} / n_{lower}	6	5	4	3	2	
1					1.634×10^{-18}	$\lvert\Delta E\rvert$ (J)
					121.6	λ (nm)
2					$\lvert\Delta E\rvert$ (J)	
					λ (nm)	
3				$\lvert\Delta E\rvert$ (J)		
				λ (nm)		
4			$\lvert\Delta E\rvert$ (J)			
			λ (nm)			
5		$\lvert\Delta E\rvert$ (J)				
		λ (nm)				

(Workshop continued on next page)

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C. Assignment of Wavelengths

Compare the wavelengths you have calculated in Table Two with those listed in Table Three. You should notice that many wavelengths match within the error of your calculation. Fill in the quantum numbers of the upper and lower states for each line whose origin you can recognize by comparison of your calculated values with the observed values. Several wavelengths will not match at all; place those in Table Four and estimate the expected $n_{hi} \rightarrow n_{lo}$ following the trends. Note that Table Two only covers transitions with n_{hi} less than or equal to six. Check your estimations by solving for ΔE and wavelengths as you did on Table Two using n_{hi} numbers greater than six.

Table Three

Wavelength	$n_{hi} \rightarrow n_{lo}$	Wavelength	$n_{hi} \rightarrow n_{lo}$	Wavelength	$n_{hi} \rightarrow n_{lo}$
97.25		410.17		1,005.0	
102.57		434.05		1,093.8	
121.57	2 → 1	486.13		1,281.8	
388.91		656.28		1,875.1	
397.01		954.62		4,050.0	

Table Four: Wavelengths you cannot assign using Table Two data

Observed Wavelength (in nm)	Predicted transition $n_{hi} \rightarrow n_{lo}$	Calculated ΔE (in J)	Calculated Wavelength λ (in nm)
388.91			

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D. The Balmer Series

When Balmer formulated his famous series for hydrogen in 1886, he was limited experimentally to wavelengths for the visible and near ultraviolet regions from 250 nm to 700 nm. All the lines in his series lie in this wavelength range. All transitions in the Balmer Series have $n_{final} = 2$.

1. What would be the longest POSSIBLE wavelength for a line in the Balmer Series?

________________nm

2. What would be the shortest POSSIBLE wavelength for a line in the Balmer Series?

________________ nm

In a normal hydrogen atom, the electron is in the *lowest* energy state. The maximum energy of an electron in the hydrogen atom is 0 J, at which point the electron is in the $n = \infty$ state, essentially removed from the atom. At this point, ionization has occurred.

3. How much energy in joules does it take to ionize the hydrogen atom?

________________ J

4. The ionization energy you have calculated is for one electron in a single hydrogen atom. Calculate the ionization energy for one mole of H atoms.

   SHOW CALCULATIONS:

________________ kJ / mole

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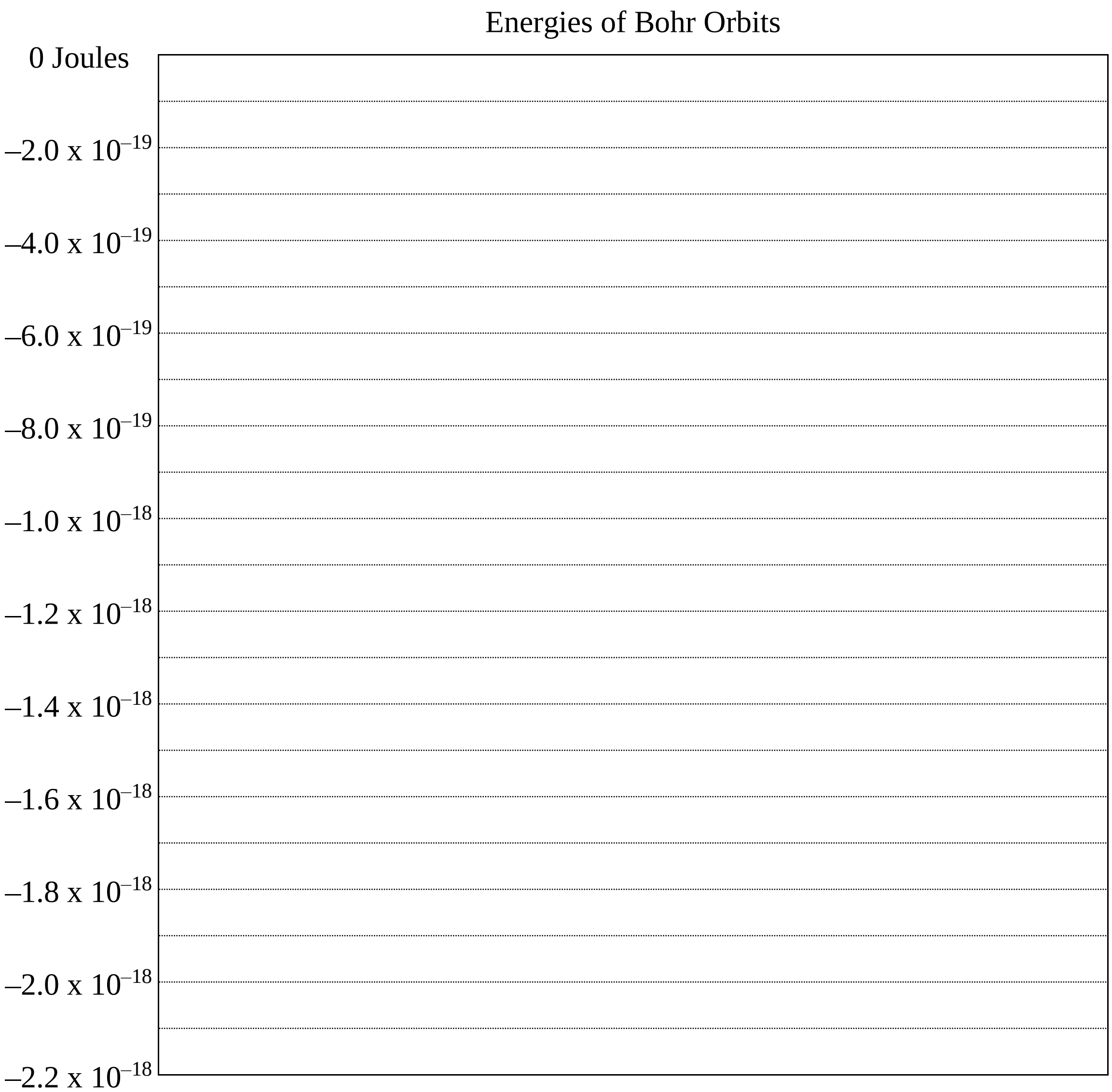
E. Energy Levels of Hydrogen Atom

Show each of the first six lowest energy states in the chart below using the values from Table One. Draw horizontal lines for each level and identify them by writing its quantum number on the right side. Use vertical arrows to show the electron transitions calculated in Table Three.



F. Hydrogen Line Spectra

Draw the line spectra for hydrogen as it would appear in the visible region showing the lines calculated in Table Three within the Balmer Series.

700 nm 250 nm

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Workshop #10: Quantum Mechanics and Chemical Periodicity

Many important facts and laws in chemistry are experimentally determined, and then rationalized in terms of a theory or artificial concept. The Periodic Law is one of these. It is based on experiment and rationalized in terms of structural concepts. This form of the Periodic Table may be explained on the basis of the order in which the electrons occupy the various energy levels. Actually, the Periodic Table is based on experiment and serves as a guide to the order in which electron-filling of shells takes place.

A relationship between the *s*, *p*, *d*, and *f* orbitals and the Periodic Table may be observed by noting that the long form of the table can be divided into blocks. One of the blocks is two elements wide, another six elements wide, a third ten elements wide, and a fourth is fourteen elements wide, respectively. Specific sections of each period and each period in the table arise from the filling of orbitals of roughly equal energy.

1. For the first problem, complete the following table for the main group elements:

Group Number	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
Number of valence electrons				4				
Electronic configuration of valence electrons. Omit principle quantum number.				s^2p^2				
Common oxidation states.				±4				

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2. For the next problem, consider the chart below, which represents the main group (representative elements) portion of the Periodic Table.

   A. Several trends in atomic properties are listed to the sides and below the chart. Convert the lines into arrows by adding arrow heads to each line to indicate the direction of each trend (i.e. → or ←).

   B. In each box, write the electronic configuration of all the valence electrons for that element. Example: see the box containing element 84 (polonium)

Nonmetallic Properties Increase

Atomic Radii Increase

IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA or 0
3	4	5	6	7	8	9	10
11	12	13	14	15	16	17	18
19	20	31	32	33	34	35	36
37	38	49	50	51	52	53	54
55	56	81	82	83	84 $6s^2 6p^4$	85	86
87	88						

Electronegativity Increases

Ionization Energy Increases

___Metallic Properties Increase___

___Atomic Radii Increase___

___Ionization Energy Increases___

___Electronegativity Increases___

(Workshop continued on next page)

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3. In each square shown below, write the principal quantum number and orbital letter of the expected last electron to enter the atom in its ground state. For this exercise, ignore the exceptions. *(Four of them have been done for you.)*

	I A 1	II A 2	III B 3	IV B 4	V B 5	VI B 6	VII B 7	---VIIIB --- 8	9	10	I B 11	II B 12	III A 13	IV A 14	V A 15	VI A 16	VII A 17	VIII A 18
1	1																	2
2	3	4											5	6	7	8	9	10
3	11	12											13	14	15	16	17	18
4	19	20	21	22	23	24	25	26	27	28	29	30 3d	31	32	33	34	35	36
5	37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52 5p	53	54
6	55	56 6s	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
7	87	88	89	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118

6	58	59	60	61	62	63	64	65	66 4f	67	68	69	70	71
7	90	91	92	93	93	95	96	97	98	99	100	101	102	103

4. A. Fill in the following table:

Quantum number *l*	0			3
Orbital Designation		*p*	*d*	

B. What m_l values are possible for the *d* orbitals?

C. What m_s values are possible?

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5. Determine the quantum numbers for all six electrons in the 4*p* sublevel.

Electron	***n***	***l***	m_l	m_s
4*p*¹				
4*p*²				
4*p*³				
4*p*⁴				
4*p*⁵				
4*p*⁶				

6. For the sets of quantum numbers below, identify its electron configuration (if possible). If not possible, explain what is wrong.

n	***l***	m_l	m_s	**electron configuration or explanation of problem**
2	0	–1	– ½	____
4	2	1	–½	____
2	0	0	+½	____
5	–1	1	0	____

7. Determine the maximum number of electrons contained in:

A. *d* sublevel ________________ B. valence (outer) shell ______________

C. a single orbital ______________ D. energy level *n* = 4 ________________

8. Write FOUR isoelectronic species for the $A1^{+3}$ ion, two cations and two anions.

9. Identify the elements which have no electron with the quantum number *l* = 1.

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10. Consider the bismuth (Bi) atom.

A. Write the complete (start with 1*s*) and shortened (noble gas in brackets) electronic configuration for bismuth. Make certain to place brackets around the closed shell (core) electrons and identify valence electrons and pseudo-core electrons.

B. Draw the orbital diagram for all of the electrons in Bi.

C. Is Bismuth paramagnetic or diamagnetic?

D. Write the set of quantum numbers describing only valence electrons in Bi.

E. Write the shortened electronic configuration for the bismuth ions below:

Bi^{+3} ion ________________________ Bi^{+5} ion ________________________

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11. A. Calculate the wavelength (in nm) of light with frequency 2.31×10^{14} Hz.

B. Visible light has wavelengths between 400 to 700 nm. Slightly longer wavelengths are infrared (IR) and shorter are ultraviolet (UV). Is electromagnetic radiation from 2.31×10^{14} Hz found to be IR, Vis, or UV?

12. A. Solve for the wavelength (in nm) caused by a hydrogen electron jumping from $n = 6$ to $n = 3$.

B. Is this photon in the visible, IR, or UV portion of the spectrum?

C. What is the frequency (in s^{-1}) for this photon?

D. Calculate the energy of this photon in both J/photon and in kJ/mol.

13. The compound known as Sunbrella, which is the active ingredient in some sunscreens, absorbs strongly around 266 nm. What is the frequency of the absorption (in MHz)?

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14. For the last problem, fill in the following table for the various chemical species

Species	Short electronic configuration	"Short" Orbital Diagram	Quantum numbers of last e$^-$	Configuration of valence electrons	Common oxidation state(s)
O	[He] $2s^2 2p^4$	[He] ↑↓ (2*s*) ↑↓ ↑ ↑ (2*p*)	2, 1, –1, ½	$2s^2\,2p^4$	–2
Si					
K					
Sr					
Cr					Varies
Mn					Varies
Ga					
As					
Mo^{+2}					N/A
Fe^{+3}					N/A
Ag^{+}					N/A

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Workshop #11: Intermolecular Forces

For the first part of this workshop, identify the type of crystal structure (Ionic, Molecular Polar, Molecular Nonpolar, Network-Covalent, or Metallic) present. Then determine the type of binding forces present in each (Ionic Bonds, Covalent Bonds, Metallic Bonds, London Dispersion Forces, Dipole Forces, and/or Hydrogen Bonds).

Substance	**Type of Crystal**	**Type of Binding Force(s)**
Ar		
CH_3Cl		
CH_3OH		
BCl_3		
CH_3OCH_3		
HF		
Hg		
KCl		
N_2		
SiC		
CH_3COOH		
Diamond		

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Circle the species with the higher boiling point and *briefly* justify your choice below.

1) Kr ____________________ or Xe ____________________
Justification:

2) C_2H_5OH ____________________ or CH_3OCH_3 ____________________
Justification:

3) NaF ____________________ or MgO ____________________
Justification:

4) N_2 ____________________ or NO ____________________
Justification:

5) CH_4 ____________________ or SiH_4 ____________________
Justification:

6) HF ____________________ or HI ____________________
Justification:

7) CO_2 ____________________ or NH_3 ____________________
Justification:

8) CH_4 ____________________ or CCl_4 ____________________
Justification:

9) Cr ____________________ or Si ____________________
Justification:

10) H_2O ____________________ or SiO_2 ____________________
Justification:

11) MgO ____________________ or BaO ____________________
Justification:

12) $CH_3CH_2CH_2CH_2CH_3$ ____________ or $(CH_3)_2CHCH_2CH_3$ ____________
Justification:

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Workshop #12: Vapor Pressure

The stronger the intermolecular forces that exist between liquid molecules, the less likely they will escape into the vapor phase. Boiling point (which you explored in Workshop #11) and vapor pressure are both good measures of intermolecular forces. In the following problem set, you will analyze some provided "experimental" data in order to calculate the vapor pressure of a liquid.

Vapor pressure is defined as the pressure of a vapor that is in equilibrium with its liquid. It is controlled by 2 factors:

1. temperature – the higher the temperature, the greater kinetic energy the liquid molecules possess; therefore, they vaporize more readily, hence increasing the vapor pressure.

2. molar heat of vaporization, ΔH_{vap} – the energy required to change a liquid to a gas at its boiling point. The stronger the intermolecular forces, the harder it is to pull liquid molecules apart, and therefore the higher its ΔH_{vap}, which decreases its vapor pressure.

The Clausius-Clapeyron Equation relates the three quantities vapor pressure, ΔH_{vap}, and temperature according to the equation:

$$\ln VP = -\frac{\Delta H_{vap}}{RT} + B$$

Notice this equation fits the slope-intercept form y = mx + b, so if ln VP is plotted against 1 / T, a straight line results with $-\Delta H_{vap}$ / R as the slope. You will use this equation and the provided "experimental" data to calculate an unknown liquid's ΔH_{vap} and its boiling point at a particular temperature. Consider the following:

	Temperature, t, (in ^{o}C)	Heights of Manometer Mercury Levels (in mm)		Vapor Pressure (in mmHg or torr)
		open to atm	atm + VP trapped on gas side	
1.	1.2	250	228	____________
2.	21.1	265	205	____________
3.	40.0	297	142	____________

4. Boiling Point: 76 ^{o}C Barometric Pressure: 752 torr

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Now fill in the following table to prepare for the graph:

t, ºC	T, Kelvin	1 / T, K^{-1}	VP, mmHg	ln VP
1.2				
21.1				
40.0				

Graph ln VP vs. 1 / T on Microsoft Office Excel® (see Experiment #2 in this lab manual for directions on using Excel®). According to the Clausius-Clapeyron equation, the slope is equal to $-\Delta H_{vap}$ / R. Using R = 8.314 x 10^{-3} kJ / mole·K, calculate ΔH_{vap} for the liquid:

slope = Δy/Δx = Δ(lnVP)/Δ(1/T) = ____________ = $-\Delta H_{vap}$/R (rearrange to solve for ΔH_{vap})

SHOW CALCULATION:

Therefore, ΔH_{vap} = ___________ kJ / mole

From the graph, you can also calculate what the liquid's boiling point should be at the "experimental" barometric pressure. Recall that boiling point is the temperature where the vapor pressure is equal to the atmospheric pressure.

"Experimental" barometric pressure ____________ mmHg
(= the VP needed for boiling)

ln (barometric pressure) ____________

1 / T at this vapor pressure ____________ K^{-1} (from the graph)

T at this pressure ____________ K

t at this pressure ____________ °C (= the predicted boiling point)

"Experimental" boiling point ____________ °C

Make sure to submit your properly labeled graph when submitting this Workshop!

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Workshop #13: Colligative Properties

Show calculation setups and answers for all problems below.

1. List the following aqueous solutions in the order of expected DECREASING FREEZING POINT: 0.075 *m* glucose; 0.075 *m* LiBr; 0.030 *m* $Zn(NO_3)_2$.

2. The normal freezing point of pure naphthalene is measured to be 80.29 °C. When 32.21 g of the nonelectrolyte urea (CH_4N_2O) is dissolved in 751.36 g of naphthalene, the freezing point is measured to be 75.34 °C. What is the molal freezing point depression constant (K_f) for naphthalene?

3. When 132.0 g of C_6H_6 (P° = 93.96 torr) and 147.0 g of $C_2H_4Cl_2$ (P° = 224.9 torr) are combined, what is the total vapor pressure of the ideal solution?

4. Calculate the freezing point of a solution of 22.0 g of carbon tetrachloride dissolved in 800.0 g of benzene (K_f = 5.12 °C / *m*; normal freezing point = 5.5 °C).

5. What mass of $NiSO_4 \cdot 6H_2O$ must be dissolved in 500. g of water to produce 0.33 *m* $NiSO_4(aq)$?

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6. What is the normal boiling point of an aqueous solution that has a freezing point of –1.04 °C?

   Note: For water, K_f = 1.86 °C/*m*; K_b = 0.512 °C/*m*

7. Assuming complete dissociation, calculate the freezing point of a 0.100 *m* aqueous solution of K_2SO_4 (ignore any interionic attractions).

   Note: For water, K_f = 1.86 °C/*m*

8. When 2.25 g of an unknown nonelectrolyte was dissolved in 150. g of cyclohexane, the boiling point increased by 0.481 K. Determine the molar mass of the compound.

   Note: K_b (cyclohexane) = 2.79 K/*m*

9. A 0.50 g sample of *immunoglobulin G*, a nonvolatile nonelectrolyte, is dissolved in enough water to make 0.100 L of solution, and the osmotic pressure of the solution at 25 °C is found to be 0.619 torr. Calculate the molecular mass of *immunoglobulin G*.

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10. When 2.74 g of phosphorus is dissolved in 100.0 mL of carbon disulfide, the boiling point is 319.71 K. Given that the normal boiling point of pure carbon disulfide is 319.30 K, its density is 1.261 g / mL, and its boiling-point elevation constant is K_b = 2.34 K / *m*, determine the molar mass of phosphorus.

11. A solution of biphenyl ($C_{12}H_{10}$), a nonvolatile nonelectrolyte, in benzene has a freezing point of 5.4 °C. Determine the osmotic pressure of the solution at 10 °C if its density is 0.88 g / cm^3.

Note: normal freezing point (benzene) = 5.5 °C; K_f = 5.12 °C/*m*

12. Consider these two solutions: Solution A is prepared by dissolving 5.00 g of $MgCl_2$ in enough water to make 0.250 L of solution, and Solution B is prepared by dissolving 5.00 g of KCl in enough water to make 0.250 L of solution. Which direction will solvent ***initially*** flow if these two solutions are separated by a semipermeable membrane?

13. Assuming that the volumes of the solutions described in question #12 are additive and ignoring any effects that gravity may have on the osmotic pressure of the solutions, what will be the ***final*** volume of solution A when the net solvent flow through the semipermeable membrane stops?