AP Chemistry
Final Exam Practice

Chapters 1-4

1) A sample of dolomitic limestone containing only $CaCO_3$ and $MgCO_3$ was analyzed.

(a) When a 0.2800 gram sample of this limestone was decomposed by heating, 75.0 milliliters of CO_2 at 750 mm Hg and 20˚C were evolved. How many grams of CO_2 were produced.

(b) Write equations for the decomposition of both carbonates described above.

(c) It was also determined that the initial sample contained 0.0448 gram of calcium. What percent of the limestone by mass was $CaCO_3$?

(d) How many grams of the magnesium-containing product were present in the sample in (a) after it had been heated?

2) An experiment is performed to determine the empirical formula of a copper iodide formed by direct combination of elements. A clean strip of copper metal is weighed accurately. It is suspended in a test tube containing iodine vapor generated by heating solid iodine. A white compound forms on the strip of copper, coating it uniformly. The strip with the adhering compound is weighed. Finally, the compound is washed completely from the surface of the metal and the clean strip is dried and reweighed.

DATA TABLE	
Mass of clean copper strip	1.2789 grams
Mass of copper strip and compound	1.2874 grams
Mass of copper strip after washing	1.2748 grams

(a) State how you would use the data above to determine each of the following. (Calculations not required.)

(1) The number of moles of iodine that reacted

(2) The number of moles of copper that reacted

(b) Explain how you would determine the empirical formula for the copper iodide.

(c) Explain how each of the following would affect the empirical formula that could be calculated.

(1) Some unreacted iodine condensed on the strip.

(2) A small amount of the white compound flaked off before weighing.

Chapter 5

3) Three volatile compounds X, Y, and Z each contain element Q. The percent by weight of element Q in each compound was determined. Some of the data obtained are given below.

Compound	Percent by Weight of Element Q	Molecular Weight
X	64.8%	?
Y	73.0%	104.
Z	59.3%	64.0

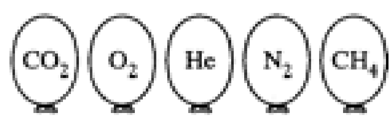
(a) The vapor density of compound X at 27 degrees Celsius and 750. mm Hg was determined to be 3.53 grams per liter. Calculate the molecular weight of compound X.

(b) Determine the mass of element Q contained in 1.00 mole of each of the three compounds.

(c) Calculate the most probable value of the atomic weight of element Q.

(d) Compound Z contains carbon, hydrogen, and element Q. When 1.00 gram of compound Z is oxidized and all of the carbon and hydrogen are converted to oxides, 1.37 grams of CO_2 and 0.281 gram of water are produced. Determine the most probable molecular formula.

4)



Represented above are five identical balloons, each filled to the same volume at 25˚C and 1.0 atmosphere 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.

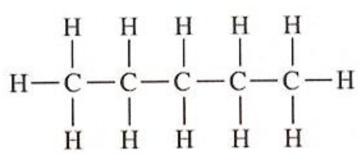
(c) Which balloon contains the gas that would be expected to deviate most from the behavior of an ideal gas? Explain.

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

Chapter 6

5) Consider the hydrocarbon pentane, C_5H_{12} (molar mass 72.15 g).

(a) Write the balanced equation for the combustion of pentane to yield carbon dioxide and water.
(b) What volume of dry carbon dioxide, measure at 25° C and 785 mm Hg, will result from the complete combustion of 2.50 g of pentane?
(c) The complete combustion of 5.00 g of pentane releases 243 kJ of heat. On the basis of this information, calculate the value of ΔH for the complete combustion of one mole of pentane.
(d) Under identical conditions, a sample of an unknown gas effuses into a vacuum at twice the rate that a sample of pentane gas effuses. Calculate the molar mass of the unknown gas.
(e) The structural formula of one isomer of pentane is shown below. Draw the structural formulas for the other two isomers of pentane. Be sure to include all atoms of hydrogen and carbon in your structures.



Chapter 7

6) Answer the following questions regarding light and its interaction with molecules, atoms, and ions.

(a) The longest wavelength of light with enough energy to break the Cl-Cl bonds in $Cl_2(g)$ is 495 nm.

(i) Calculate the frequency, in s^{-1}, of the light.

(ii) Calculate the energy, in J, of a photon of the light.

(iii) Calculate the minimum energy, in kJ/mol, of the Cl-Cl bond.

(b) A certain line in the spectrum of atomic hydrogen is associated with the electronic transition in the H atom from the sixth energy level (n=6) to the second energy level (n=2).

(i) Indicate whether the H atom emits energy or whether it absorbs energy during the transition. Justify your answer.

(ii) Calculate the wavelength, in nm, of the radiation associated with the spectral line.

(iii) Account for the observation that the amount of energy associated with the same electronic transition (n=6 to n=2) in the He^+ ion is greater than that associated with the corresponding transition in the H atom.

7) The emission spectrum of hydrogen consists of several series of sharp emission lines in the ultraviolet (Lyman series) in the visible (Balmer series) and in the infrared (Paschen series, Brackett series, etc.) regions of the spectrum.

(a) What feature of the electronic energies of the hydrogen atom explains why the emission spectrum consists of discrete wavelength rather than a continuum wavelength?

(b) Account for the existence of several series of lines in the spectrum. What quantity distinguishes one series of lines from another?

(c) Draw an electronic energy level diagram for the hydrogen atom and indicate on it the transition corresponding to the line of lowest frequency in the Balmer series.

(d) What is the difference between an emission spectrum and an absorption spectrum? Explain why the absorption spectrum of atomic hydrogen at room temperature has only the lines of the Lyman series.

8) Use the details of modern atomic theory to explain each of the following experimental observations.

(a) Within a family such as the alkali metals, the ionic radius increases as the atomic number increases.

(b) The radius of the chlorine atom is smaller than the radius of the chloride ion, Cl^-. (Radii : Cl atom = 0.99Å; Cl^- ion = 1.81Å)

(c) The first ionization energy of aluminum is lower than the first ionization energy of magnesium. (First ionization energies: $_{12}Mg$ = 7.6 ev; $_{13}Al$ = 6.0 ev)

(d) For magnesium, the difference between the second and third ionization energies is much larger than the difference between the first and second ionization energies. (Ionization energies for Mg: 1^{st} = 7.6 ev; 2^{nd} = 14 ev; 3^{rd} = 80 ev)

Chapter 8-9

9) Using principles of chemical bonding and/or intermolecular forces, explain each of the following.

(a) Xenon has a higher boiling point than neon has.

(b) Solid copper is an excellent conductor of electricity, but solid copper chloride is not.

(c) SiO_2 melts at a very high temperature, while CO_2 is a gas at room temperature, even though Si and C are in the same chemical family.

(d) Molecules of NF_3 are polar, but those of BF_3 are not.

10)

CF_4 XeF_4 ClF_3

(a) Draw a Lewis electron–dot structure for each of the molecules above and identify the shape of each.

(b) Use the valence shell electron–pair repulsion (VSEPR) model to explain the geometry of each of these molecules.

11) Use simple structure and bonding models to account for each of the following.

(a) The bond length between the two carbon atoms is shorter in C_2H_4 than in C_2H_6.

(b) The H-N-H bond angle is 107.5°, in NH_3.

(c) The bond lengths in SO_3 are all identical and are shorter than a sulfur-oxygen single bond.

(d) The I_3^- ion is linear.

Chapter 10

12) The normal boiling and freezing points of argon are 87.3 K and 84.0 K, respectively. The triple point is at 82.7 K and 0.68 atmosphere.

(a) Use the data above to draw a phase diagram for argon. Label the axes and label the regions in which the solid, liquid and gas phases are stable. On the phase diagram, show the position of the normal boiling point.
(b) Describe any changes that can be observed in a sample of solid argon when the temperature is increased from 40 K to 160 K at a constant pressure of 0.50 atmosphere.
(c) Describe any changes that can be observed in a sample of liquid argon when the pressure is reduced from 10 atmospheres to 1 atmosphere at a constant temperature of 100 K, which is well below the critical temperature.
(d) Does the liquid phase of argon have a density greater than, equal to, or less than the density of the solid phase? Explain your answer, using information given in the introduction to this question.

13) Explain each of the following in terms of atomic and molecular structures and/or intermolecular forces.
(a) Solid K conducts an electric current, whereas solid KNO_3 does not.
(b) $SbCl_3$ has measurable dipole moment, whereas $SbCl_5$ does not.
(c) The normal boiling point of CCl_4 is 77˚C, whereas that of CBr_4 is 190˚C.
(d) $NaI_{(s)}$ is very soluble in water, whereas $I_{2(s)}$ has a solubility of only 0.03 gram per 100 grams of water.

14) For each of the following, use appropriate chemical principles to explain the observation.
(a) Sodium chloride may be spread on an icy sidewalk in order to melt the ice; equimolar amounts of calcium chloride are even more effective.
(b) At room temperature, NH_3 is a gas and H_2O is a liquid, even though NH_3 has a molar mass of 17 grams and H_2O has a molar mass of 18 grams.
(c) C (graphite) is used as a lubricant, whereas C (diamond) is used as an abrasive.
(d) Pouring vinegar onto the white residue inside a kettle used for boiling water results in a fizzing/bubbling phenomenon.

Answers:

1)

(a)

$= 3.08\times10^{-3}$ mol

$$3.08\times10^{-3}\text{ mol} \times \frac{44.0\text{ g } CO_2}{1\text{ mol}} = 0.135\text{ g } CO_2$$

(b) $CaCO_3 \rightarrow CaO + CO_2$

$MgCO_3 \rightarrow MgO + CO_2$

(c)

$$\frac{0.112\text{ g } CaCO_3}{0.2800\text{ g sample}} = 40.0\%\ CaCO_3$$

(d) 60.0% of 0.2800 g sample = 0.168 g of $MgCO_3$

2) (a) (1) (mass of Cu Strip + compound) - (mass of original clean Cu strip) = mass of iodine

(mass of iodine)/(atomic mass of iodine) = moles of iodine in the sample of compound

(2) (mass of original clean Cu strip) - (mass of strip after washing and drying) = mass of Cu

(mass of Cu) / (atomic mass of Cu) = moles of Cu in sample of compound

(b) The empirical formula is the ratio (moles iodine) / (moles Cu). **OR** (moles Cu) / (moles iodine).

(c) (1) Unreacted I_2 would make the apparent mass of compound and the iodine too high. Thus, the I:Cu ratio in the empirical formula would be too high.

(2) If some compound flaked off, the mass of compound (and the I_2) would be too low. Thus the I:Cu ratio in the empirical formula would be too low.

3)

(a)

=88.1g/mol

(b)

	X	Y	Z
	88.1 g/mol	104	64.0
% Q	64.8	73.0	59.3
g Q	57.1	75.9	38.0

(c) ratio 1.5 2 1

masses must be integral multiples of atomic weight

therefore, 3 4 2

which gives an atomic weight of Q = 19

(d)

1.00 g Z is 59.3% Q = 0.593 g Q

$$0.0593 \text{ g Q} \times \frac{1 \text{ mol}}{19.0 \text{ g}} = 0.0312 \text{ mol Q}$$

therefore, the empirical formula = CHQ, the smallest whole number ratio of moles.

formula wt. of CHQ = 32.0, if mol. wt. Z = 64 then the formula of Z = $(CHQ)_2$ or $C_2H_2Q_2$

4) (a) CO_2; according to Avogadro's Hypothesis, they all contain the same number of particles, therefore, the heaviest molecule, CO_2 (molar mass = 44), will have the greatest mass.

(b) all the same; at the same temperature all gases have the same kinetic energy.

(c) CO_2; since they are all essentially non-polar, the largest intermolecular (London) force would be greatest in the molecule/atom with the largest number of electrons.

(d) He; it has the smallest size and has the greatest particulate speed and, therefore, it's the easiest to penetrate the wall and effuse.

5) a) $C_5H_{12} + 8O_2 = 5CO_2 + 6H_2O$

b) 4.10 L

c) — 3506 kj/mole

d) 18 g/mole

e) 2 methyl butane and 2,2, di methyl propane

6) No answer available

7) (a) Any of the following:

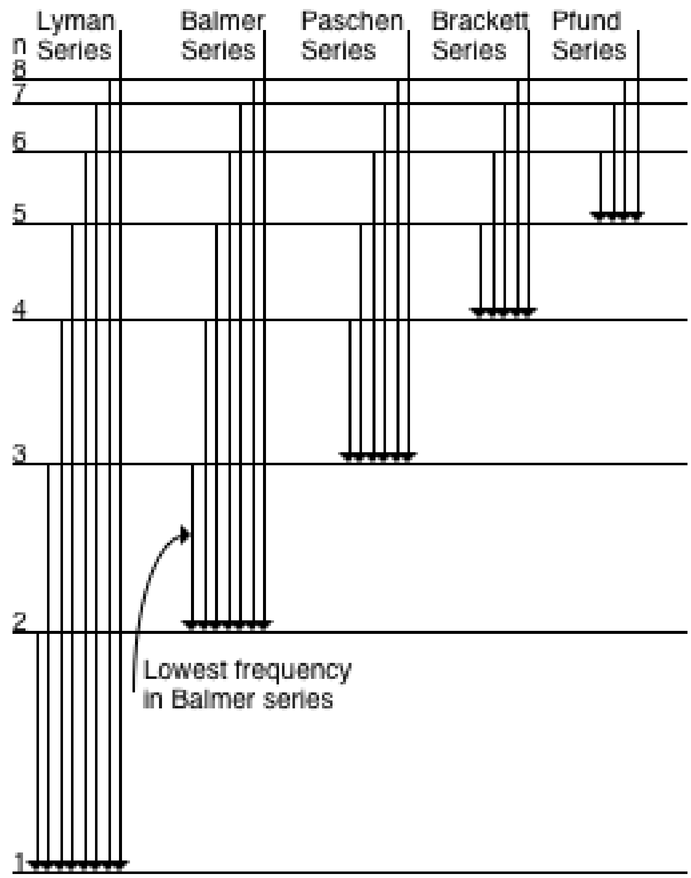
Quantized energy levels. Discrete energies.

Wave properties of electron result in discrete energy state.

(b) An electron in an excited-state atom can go to any of several lower energy states.

The lines in each series represents shifts from several higher energy states to a single lower energy state, identified by the same principal quantum number or energy.

(c)



(d) Emission spectra obtained when electrons in excited atoms drop to lower energy levels.

Absorption spectra obtained when electrons in atoms in ground (or lower energy) state absorb electromagnetic radiation and move to higher energy states.

H atoms at 25˚C are in lowest electronic energy state (n = 1) and so the only absorptions will result from electrons moving from n = 1 to higher levels.

8) (a) The radii of the alkali metal ions increase with increasing atomic number because the outer principal quantum number (or shell or energy level) is larger. **OR**

(1) There is an increase in shielding. (2) The number of orbitals increases.

(b) The chloride ion is larger than the chlorine atom because - (any of these)

(1) the electron-electron repulsion increases.

(2) the electron-proton ratio increases.

(3) the effective nuclear charge decreases.

(4) shielding increases.

(c) The first ionization energy for Mg is greater than that for Al because - *(either of these)*

(1) the 3*p* orbital (Al) represents more energy than the 3*s* orbital (Mg) represents.

(2) the 3*p* electron in an Al atom is better shielded from its nucleus than a 3*s* electron in a Mg atom.

(3) [half credit] a 3*p* electron is easier to remove than a 3*s* electron.

(d) In a Mg atom, the first two electrons lost are removed from the 3*s* orbital whereas the 3rd electron comes from a 2*p* orbital; a 2*p* orbital is much lower in energy than the 3*s* is; so more energy is needed to remove a 2*p* electron.

9)
Answer:

(a) Xe and Ne are monatomic elements held together by London dispersion (van der Waals) forces. The magnitude of such forces is determined by the number of electrons in the atom. A Xe atom has more electrons than a neon atom has. (Size of the atom was accepted but mass was not.)

(b) The electrical conductivity of copper metal is based on mobile valence electrons (partially filled bands). Copper chloride is a rigid ionic solid with the valence electrons of copper localized in individual copper(II) ions.

(c) SiO_2 is a covalent network solid. There are strong bonds, many of which must be broken simultaneously to volatilize SiO_2. CO_2 is composed of discrete, nonpolar CO_2 molecules so that the only forces holding the molecules together are the weak London dispersion (van der Waals) forces.

(d) In NF_3 a lone pair of electrons on the central atom results in a pyramidal shape. The dipoles don't cancel, thus the molecule is polar.

While in BF_3 there is no lone pair on the central atom so the molecule has a trigonal planar shape in which the dipoles cancel, thus the molecule is nonpolar.

10) (a)

(b) CF_4 = 4 bonding pairs around C at corners of regular tetrahedron to minimize repulsion (maximize bond angles).

XeF_4 = 4 bonding pairs and 2 lone pairs give octahedral shape with lone pairs on opposite sides of Xe atom

ClF_3 = 3 bonding pairs and 2 lone pairs give trigonal bipyramid with one pairs in equatorial positions 120º apart.

11) (a) C_2H_4 has a multiple bond; C_2H_6 has a single bond. Multiple bonds are stronger and, therefore, shorter than single bonds.

(b) NH_3 has 3 bonding pairs of electrons and 1 lone pair. Bonding pairs are forced together because repulsion between lone pair and bonding pairs is greater than between bonding pairs.

(c) The bonding in SO_3 can be described as a combination of 3 resonance forms of 1 double and 2 single bonds.

The actual structure is intermediate among the 3 resonance forms, having 3 bonds that are equal and stronger (therefore, shorter) than an S-O single bond.

(d) The central I atom has 3 lone pairs and 2 bonding pairs around it.
To minimize repulsion, the 3 lone pairs on the central atom are arranged as a triangle in a plane are right angles to the I-I-I- axis.

12)

(a)

(b) The argon sublimes.

(c) The argon vaporizes.

(d) The liquid phase is less dense than the solid phase. Since the freezing point of argon is higher than the triple point temperature, the solid-liquid equilibrium line slopes to the right with increasing pressure. Thus, if a sample of liquid argon is compressed (pressure increased) at constant temperature, the liquid becomes a solid. Because increasing pressure favors the denser phase, solid argon must be the denser phase.

13) (a) K conducts because of its metallic bonding - or – "sea" of mobile electrons (or free electrons). KNO_3 does not conduct because it is ionically bonded and has immobile ions (or immobile electrons).

(b) $SbCl_3$ has a measurable dipole moment because it has a lone pair of electrons which causes a dipole - or - its dipoles do not cancel - or - it has a trigonal pyramidal structure - or - a clear diagram illustrating any of the above.

(c) CBr_4 boils at a higher temperature than CCl_4 because it has stronger intermolecular forces (or van der Waal or dispersion). These stronger forces occur because CBr_4 is larger and/or has more electrons than CCl_4.

(d) NaI has greater aqueous solubility than I_2 because NaI is ionic (or polar), whereas I_2 is non-polar (or covalent). Water, being polar, interacts with the ions of NaI but not with I_2. (Like dissolves like accepted if polarity of water is clearly indicated.)

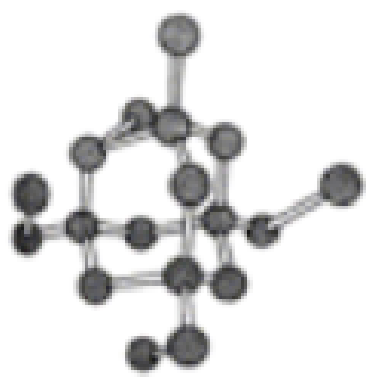
14) (a) $NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

$CaCl_{2(s)} \rightarrow Ca^+_{(aq)} + 2\ Cl^-_{(aq)}$

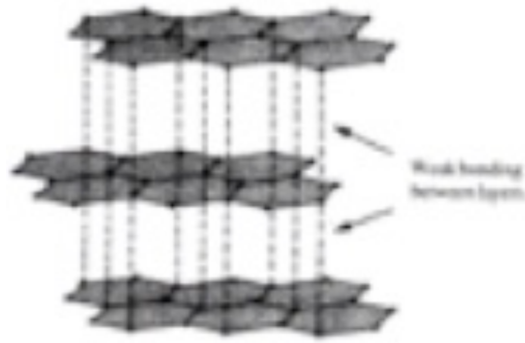
The freezing point of an aqueous solution is lower than the freezing point of water. A higher molality of a solution lowers the freezing point more and an equimolar amount of the two solids gives a larger molal solution from the calcium chloride as illustrated by the above equations.

(b) Water is more polar than ammonia creating stronger attractions (IMF) between molecules and making it a liquid.

(c) Diamond, the hardest naturally occurring substance, has each carbon atom surrounded by a tetrahedral arrangement of other carbon atoms (see drawing). The network solid structure is stabilized by covalent bonds, formed by the overlap of *sp^3* hybridized carbon atomic orbitals. A diamond has uniform very strong bonds in all directions in the crystal.



Graphite has a different kind of bonding based on layers of carbon atoms arranged in fused six-member rings (see drawing). Each carbon atom in a layer is surrounded by three other carbons in a trigonal planar arrangement with 120° bond angles. The slipperiness is caused by noting that graphite has very strong bonds within layers but little bonding between the layers which allows the layers to slide past one another quite readily.



(d) Calcium and magnesium carbonates are left behind from the evaporation of hard water. These carbonates decompose and release carbon dioxide gas when reacted with the acetic acid in the vinegar.

$CaCO_{3(s)} + MgCO_{3(s)} + CH_3COOH_{(aq)} \rightarrow Ca^{2+}_{(aq)} + Mg^{2+}_{(aq)} + H_2O_{(l)} + CO_{2(g)} + CH_3COO^-_{(aq)}$