Name ________________________ **AP CHEM** ___/___/___

Collected AP Exam Essays for Chapters 7 - 10 Exam

1980 - #9

(a) Write the ground-state electron configuration for an arsenic atom, showing the number of electrons in each subshell.
(b) Give one permissible set of four quantum numbers for each of the outermost electrons in a single As atom when it is in its ground state.
(c) Is an isolated arsenic atom in the ground state paramagnetic or diamagnetic? Explain briefly.
(d) Explain how the electron configuration of the arsenic atom in the ground state is consistent with the existence of the following known compounds: Na_3As, $AsCl_3$, and AsF_5.

1982 - #6

The values of the first three ionization energies (I_1, I_2, I_3) for magnesium and argon [in kJ/mole] are as follows:

	I_1	I_2	I_3
Mg	735	1443	7730
Ar	1525	2665	3945

(a) Give the electronic configuration of Mg and Ar.
(b) In terms of these configurations, explain why the values of the first and second ionization energies of Mg are significantly lower than the values for Ar, whereas the third ionization energy of Mg is much larger than the third ionization energy Ar.
(c) If a sample of Ar in one container and a sample of Mg in another container are each heated and chlorine is passed in to each container, what compounds, if any, will be formed? Explain in terms of the electronic configuration given in part (a).
(d) Element Q has the following first three ionization energies [in kJ/mole]:

	I_1	I_2	I_3
Q	496	4568	6920

What is the formula for the most likely compound of element Q with chlorine? Explain the choice of formula on the basis of the ionization energies.

1982 - #9

(a) Draw the Lewis electron-dot structures for CO_3^{2-}, CO_2, and CO, including resonance structures where appropriate.
(b) Which of the three species has the shortest C-O bond length? Explain the reason for your answer.
(c) Predict the molecular shapes for the three species. Explain how you arrived at your predictions.

1984 - #8

Discuss some differences in physical and chemical properties of metals and nonmetals. What characteristic of the electronic configurations of atoms distinguishes metals from nonmetals? On the basis of this characteristic, explain why there are many more metals than nonmetals.

1985 - #9

Substance	Melting Point, °C
H_2	-259
C_3H_8	-190
HF	-92
CsI	621
LiF	870
SiC	>2,000

(a) Discuss how the trend in the melting points of the substances tabulated above can be explained in terms of the types of attractive forces and/or bonds in these substances.
(b) For any pairs of substances that have the same kind(s) of attractive forces and/or bonds, discuss the factors that cause variation in the strengths of the forces and/or bonds.

1987 - #5

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 = 0.99 Å; Cl^- ion = 1.81 Å)
(c) The first ionization energy of aluminum is lower than the first ionization energy of magnesium. (First ionization energies: 12Mg = 7.6 ev, 13Al = 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, in electron-volts, for Mg: 1st = 7.6, 2nd = 14, 3rd = 80)

1987 - #9

Two important concepts that relate to the behavior of electrons in atomic system are the Heisenberg uncertainty principle and the wave-particle duality of matter.
(a) State the Heisenberg uncertainty principle as it relates to determining the position and momentum of an object.
(b) What aspect of the Bohr theory of the atom is considered unsatisfactory as a result of the Heisenberg uncertainty principle?
(c) Explain why the uncertainty principle or the wave nature of particles is not significant when describing the behavior of macroscopic objects, but is very significant when describing the behavior of electrons.

1988 - #5

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, event though Si and C are in the same chemical family.
(d) Molecules of NF_3 are polar, but those of BF_3 are not.

1988 - #8

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 increases from 40 K to 160 K at a constant pressure of 0.50 atmospheres.
(c) Describe any changes that can be observed in a sample of liquid argon 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.

1989 - #5

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.

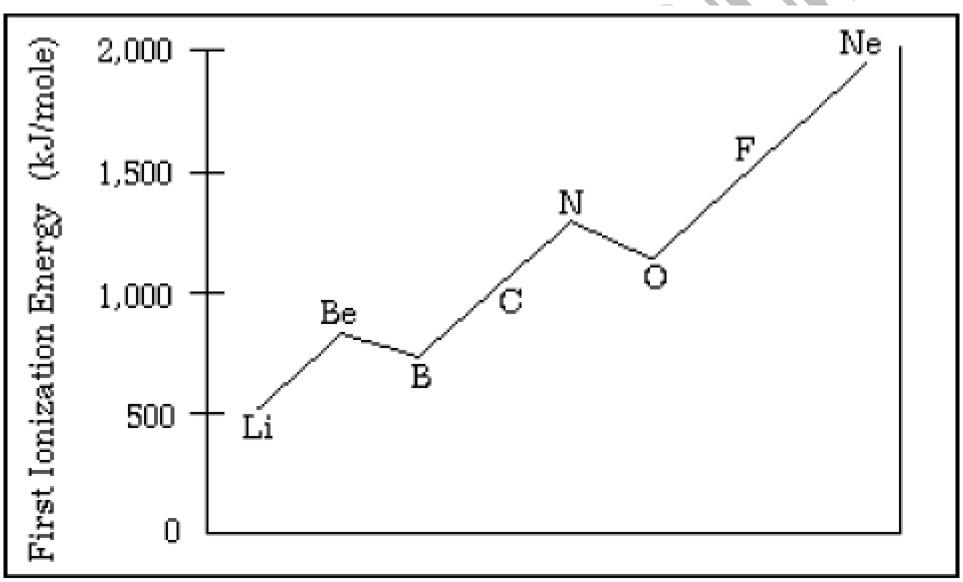
1989 - #6

The melting points of the alkali metals decrease from Li to Cs. In contrast, the melting of the halogens increases from F_2 to I_2.
(a) Using bonding principles, account for the decrease in the melting point of the alkali metals.
(b) Using bonding principles, account for the increase in the melting points of the halogens.
(c) What is the expected trend in the melting points of the compounds LiF, NaCl, KBr, and CsI? Explain this trend using bonding principles.

1990 - #5

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.

1990 - #6



The diagram shows the first ionization energies for the elements from Li to Ne. Briefly (in one to three sentences) explain each of the following in terms of atomic structure.
(a) In general, there is an increase in the first ionization energy from Li to Ne.
(b) The first ionization energy of B is lower than that of Be.
(c) The first ionization energy of O is lower than that of N.
(d) Predict how the first ionization energy of Na compares to those of Li and of Ne. Explain.

1992 - #8

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 a 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.

1992 - #9

NO_2 **NO_2^-** **NO_2^+**

Nitrogen is the central atom in each of the species given above.
(a) Draw the Lewis electron-dot structure for each of the three species.
(b) List the species in order of increasing bond angle. Justify your answer.
(c) Select one of the species and give the hybridization of the nitrogen atom in it.
(d) Identify the only one of the species that dimerizes and explain what causes it to do so.

1993 - #6

Account for each of the following in terms of principles of atomic structure, including the number, properties, and arrangements of subatomic particles.
(a) The second ionization energy of sodium is about three times greater than the second ionization energy of magnesium.
(b) The difference between the atomic radii of Na and K is relatively large compared to the difference between the atomic radii of Rb and Cs.
(c) A sample of solid nickel chloride is attracted into a magnetic field, whereas a sample of solid zinc chloride is not.
(d) Phosphorus forms the fluorides PF_3 and PF_5, whereas nitrogen forms only NF_3.

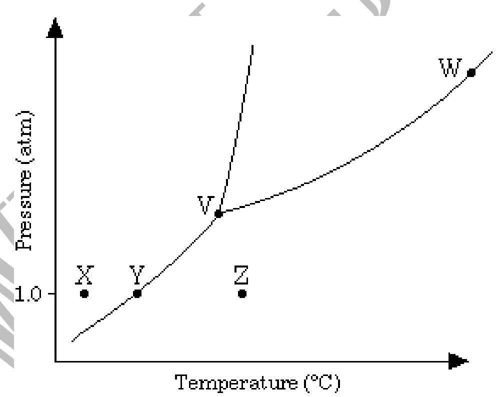
1994 - #9

Use principle of atomic structure and/or chemical bonding to answer of each of the following.
(a) The radius of the Ca atom is 0.197 nanometer; the radius of the Ca^{2+} ion is 0.099 nanometer.
Account for this difference.
(b) The lattice energy of CaO(s) is -3,460 kilojoules per mole; the lattice energy for K_2O(s) is -2,240 kilojoules per mole. Account for this difference.

Ionization Energy (kJ/mole)		
	First	Second
K	419	3050
Ca	590	1140

(c) Explain the difference between Ca and K in regard to:
(i) their first ionization energies.
(ii) their second ionization energies.
(d) The first ionization energy of Mg is 738 kilojoules per mole and that of Al is 578 kilojoules per mole. Account for this difference.

1995 - #6



The phase diagram for a pure substance is shown above. Use this diagram and your knowledge about changes of phase to answer the following questions.
(a) What does point V represent? What characteristics are specific to the system only at point V?
(b) What does each point on the curve between V and W represent?
(c) Describe the changes that the system undergoes as the temperature slowly increases from X to Y to Z at 1.0 atmosphere.
(d) In a solid-liquid mixture of this substance, will the solid float or sink? Explain.

1995 - #7

Explain the following in terms of the electronic structure and bonding of the compounds considered.

(a) Liquid oxygen is attracted to a strong magnet, whereas liquid nitrogen is not.

(b) The SO_2 molecule has a dipole moment, whereas the CO_2 molecule has no dipole moment. Include the Lewis (electron-dot) structures in your explanation.

(c) Halides of cobalt(II) are colored, whereas halides of zinc(II) are colorless.

(d) A crystal of high purity silicon is a poor conductor of electricity; however, the conductivity increases when a small amount of arsenic is incorporated (doped) into the crystal.

1996 - #9

Explain each of the following in terms of the electronic structure and/or bonding of the compounds involved.

(a) At ordinary conditions, HF (normal boiling point = 20°C) is a liquid, whereas HCl (normal boiling point = -114°C) is a gas.

(b) Molecules of AsF_3 are polar, whereas molecules of AsF_5 are nonpolar.

(c) The N-O bonds in the NO_2^- ion are equal in length, whereas they are unequal in HNO_2.

(d) For sulfur, the fluorides SF_2, SF_4, and SF_6 are known to exist, whereas for oxygen only OF_2 is known to exist.

1997 - #5

Consider the molecules PF_3 and PF_5.

(a) Draw the Lewis electron-dot structures for PF_3 and PF_5 and predict the molecular geometry of each.

(b) Is the PF_3 molecular polar, or is it nonpolar? Explain.

(c) On the basis of bonding principles, predict whether each of the following compounds exists. In each case, explain your prediction.

(i) NF_5

(ii) AsF_5

1997 - #6

Explain each of the following observations using principles of atomic structure and/or bonding.

(a) Potassium has a lower first-ionization energy than lithium.

(b) The ionic radius of N^{3-} is larger than that of O^{2-}.

(c) A calcium atom is larger than a zinc atom.

(d) Boron has a lower first-ionization energy than beryllium.

1999 - #2

Answer the following questions regarding light and its interactions with molecules, atoms, and ions.

(a) The longest wavelength of light with enough energy to break the Cl-Cl bond 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^{-1}, 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.

1999 - #8

Answer the following questions using principles of chemical bonding and molecular structure.

Consider the carbon dioxide molecule, CO_2 , and the carbonate ion, CO_3^{2-}.

a. Draw the complete Lewis electron-dot structure for each species.

b. Account for the fact that the carbon-oxygen bond length in CO_3^{2-} is greater than the carbon-oxygen bond length in CO_2.

Consider the molecules CF_4 and SF_4.

a. Draw the complete Lewis electron-dot structure for each molecule.

b. In terms of molecular geometry, account for the fact that the CF_4 molecule is nonpolar, whereas the SF_4 molecule is polar.

2000 - #7
Answer the following questions about the element selenium, Se (atomic number 34).
(a) Samples of natural selenium contain six stable isotopes. In terms of atomic structure, explain what these isotopes have in common, and how they differ.
(b) Write the complete electron configuration (e.g., $1s^2\ 2s^2$. . . etc.) for a selenium atom in the ground state. Indicate the number of unpaired electrons in the ground-state atom, and explain your reasoning.
(c) In terms of atomic structure, explain why the first ionization energy of selenium is
(i) less than that of bromine (atomic number 35), and
(ii) greater than that of tellurium (atomic number 52).
(d) Selenium reacts with fluorine to form SeF_4. Draw the complete Lewis electron-dot structure for SeF_4 and sketch the molecular structure. Indicate whether the molecule is polar or nonpolar, and justify your answer.

2001 - #8
Account for each of the following observations about pairs of substances. In your answers, use appropriate principles of chemical bonding and/or intermolecular forces. In each part, your answer must include references to both substances.
(a) Even though NH_3 and CH_4 have similar molecular masses, NH_3 has a much higher normal boiling point (-33°C) than CH_4 (-164°C).
(b) At 25°C and 1.0 atm, ethane (C_2H_6) is a gas and hexane (C_6H_{14}) is a liquid.
(c) Si melts at a much higher temperature (1,410°C) than Cl_2 (-101°C).
(d) MgO melts at a much higher temperature (2,852°C) than NaF (993°C).

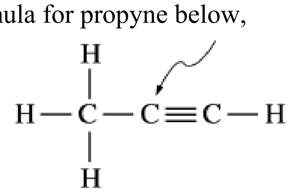
2002 - #6
Use the principles of atomic structure and/or chemical bonding to explain each of the following. In each part, your answer must include references to both substances.
(a) The atomic radius of Li is larger than that of Be.
(b) The second ionization energy of K is greater than the second ionization energy of Ca.
(c) The carbon-to-carbon bond energy in C_2H_4 is greater than it is in C_2H_6.
(d) The boiling point of Cl_2 is lower than the boiling point of Br_2.

2002B - #6
Using principles of chemical bonding and molecular geometry, explain each of the following observations. Lewis electron-dot diagrams and sketches of molecules may be helpful as part of your explanations. For each observation, your answer must include references to both substances.
(a) The bonds in nitrite ion, NO_2^-, are shorter than the bonds in nitrate ion, NO_3^-.
(b) The CH_2F_2 molecule is polar, whereas the CF_4 molecule is not.
(c) The atoms in a C_2H_4 molecule are located in a single plane, whereas those in a C_2H_6 molecule are not.
(d) The shape of a PF_5 molecule differs from that of an IF_5 molecule.
(e) $HClO_3$ is a stronger acid than HClO.

2003 - #8 a & d

Compound Name	Compound Formula
Propane	$CH_3CH_2CH_3$
Propanone	CH_3COCH_3
1-propanol	$CH_3CH_2CH_2OH$

Using the information in the table above, answer the following questions about organic compounds.
(a) For propanone,
(i) draw the complete structural formula (showing all atoms and bonds);
(ii) predict the approximate carbon-to-carbon-to-carbon bond angle.
(d) Given the structural formula for propyne below,



(i) indicate the hybridization of the carbon atom indicated by the arrow in the structure above;
(ii) indicate the total number of sigma (σ) bonds and the total number of pi (π) bonds in the molecule.

2003B - #7
Account for the following observations using principles of atomic structure and/or chemical bonding. In each part, your answer must include specific information about both substances.
(a) The Ca^{2+} and Cl^{-} ions are isoelectronic, but their radii are not the same. Which ion has the larger radius? Explain.
(b) Carbon and lead are in the same group of elements, but carbon is classified as a nonmetal and lead is classified as a metal.
(c) Compounds containing Kr have been synthesized, but there are no known compounds that contain He.
(d) The first ionization energy of Be is 900 kJ mol^{-1}, but the first ionization energy of B is 800 kJ mol^{-1}.

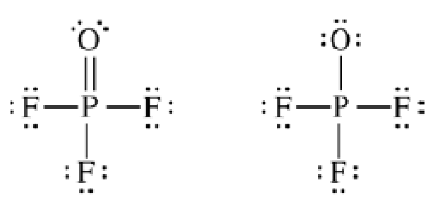
2004 - #7
Use appropriate chemical principles to account for each of the following observations. In each part, your response must include specific information about both substances.
(a) At 25°C and 1 atm, F_2 is a gas, whereas I_2 is a solid.
(b) The melting point of NaF is 993°C, whereas the melting point of CsCl is 645°C.
(c) The shape of the ICl_4^- ion is square planar, whereas the shape of the BF_4^- ion is tetrahedral.
(d) Ammonia, NH_3, is very soluble in water, whereas phosphine, PH_3, is only moderately soluble in water.

2004 - #8 a & b
Answer the following questions about carbon monoxide, CO(*g*), and carbon dioxide, CO_2(*g*). Assume that both gases exhibit ideal behavior.
(a) Draw the complete Lewis structure (electron-dot diagram) for the CO molecule and for the CO_2 molecule.
(b) Identify the shape of the CO_2 molecule.

2005 - #6
Answer the following questions that relate to chemical bonding.
(a) In the boxes provided, draw the complete Lewis structure (electron-dot diagram) for each of the three molecules represented below.

CF_4	PF_5	SF_4

(b) On the basis of the Lewis structures drawn above, answer the following questions about the particular molecule indicated.
(i) What is the F –C –F bond angle in CF_4?
(ii) What is the hybridization of the valence orbitals of P in PF_5?
(iii) What is the geometric shape formed by the atoms in SF_4?
(c) Two Lewis structures can be drawn for the OPF_3 molecule, as shown below.



Structure 1 Structure 2

(i) How many sigma bonds and how many pi bonds are in structure 1?
(ii) Which one of the two structures best represents a molecule of OPF_3? Justify your answer in terms of formal charge.

2005 - #7 a, b, & c

Use principles of atomic structure, bonding, and/or intermolecular forces to respond to each of the following. Your responses must include specific information about all substances referred to in each question.

(a) At a pressure of 1 atm, the boiling point of $NH_3(l)$ is 240 K, whereas the boiling point of $NF_3(l)$ is 144 K.
(i) Identify the intermolecular force(s) in each substance.
(ii) Account for the difference in the boiling points of the substances.

(b) The melting point of KCl(*s*) is 776°C, whereas the melting point of NaCl(*s*) is 801°C.
(i) Identify the type of bonding in each substance.
(ii) Account for the difference in the melting points of the substances.

(c) As shown in the table below, the first ionization energies of Si, P, and Cl show a trend.

Element	**First Ionization Energy ($kJ\ mol^{-1}$)**
Si	786
P	1,012
Cl	1,251

(i) For each of the three elements, identify the quantum level (e.g., $n = 1$, $n = 2$, etc.) of the valence electrons in the atom.
(ii) Explain the reasons for the trend in first ionization energies.

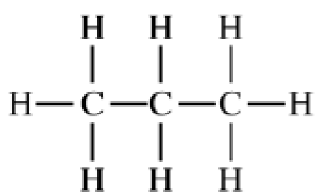
2005B - #8

Use principles of atomic structure, bonding, and intermolecular forces to answer the following questions. Your responses must include specific information about all substances referred to in each part.

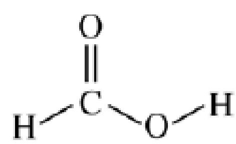
(a) Draw a complete Lewis electron-dot structure for the CS_2 molecule. Include all valence electrons in your structure.

(b) The carbon-to-sulfur bond length in CS_2 is 160 picometers. Is the carbon-to-selenium bond length in CSe_2 expected to be greater than, less than, or equal to this value? Justify your answer.

(c) The bond energy of the carbon-to-sulfur bond in CS_2 is 577 kJ mol^{-1}. Is the bond energy of the carbon-to selenium bond in CSe_2 expected to be greater than, less than, or equal to this value? Justify your answer.



Propane



Methanoic Acid

(d) The complete structural formulas of propane, C_3H_8, and methanoic acid, HCOOH, are shown above. In the table below, write the type(s) of intermolecular attractive force(s) that occur in each substance.

Substance	Boiling Point	Intermolecular Attractive Force(s)
Propane	229 K	
Methanoic acid	374 K	

(e) Use principles of intermolecular attractive forces to explain why methanoic acid has a higher boiling point than propane.

2006 - #6 a, b & c

Answer each of the following in terms of principles of molecular behavior and chemical concepts.

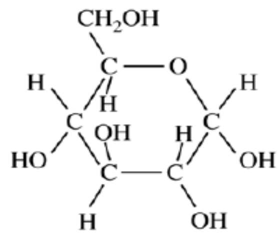
(a) The structures for glucose, $C_6H_{12}O_6$, and cyclohexane, C_6H_{12}, are shown below.

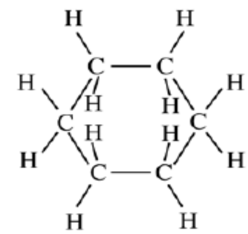
Identify the type(s) of intermolecular attractive forces in

(i) pure glucose

(ii) pure cyclohexane

(b) Glucose is soluble in water but cyclohexane is not soluble in water. Explain.





(c) Consider the two processes represented below.

Process 1: $H_2O(l) \rightarrow H_2O(g)$ $\Delta H° = +\ 44.0$ kJ mol^{-1}

Process 2: $H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g)$ $\Delta H° = +\ 286$ kJ mol^{-1}

(i) For each of the two processes, identify the type(s) of intermolecular or intramolecular attractive forces that must be overcome for the process to occur.

(ii) Indicate whether you agree or disagree with the statement in the box below. Support your answer with a short explanation.

When water boils, H_2O molecules break apart to form hydrogen molecules and oxygen molecules.

2006 - #7

Answer the following questions about the structures of ions that contain only sulfur and fluorine.

(a) The compounds SF_4 and BF_3 react to form an ionic compound according to the following equation.

$SF_4 + BF_3 \rightarrow SF_3BF_4$

(i) Draw a complete Lewis structure for the SF_3^+ cation in SF_3BF_4.

(ii) Identify the type of hybridization exhibited by sulfur in the SF_3^+ cation.

(iii) Identify the geometry of the SF_3^+ cation that is consistent with the Lewis structure drawn in part (a)(i).

(iv) Predict whether the F–S–F bond angle in the SF_3^+ cation is larger than, equal to, or smaller than 109.5°. Justify your answer.

(b) The compounds SF_4 and CsF react to form an ionic compound according to the following equation.

$SF_4 + CsF \rightarrow CsSF_5$

(i) Draw a complete Lewis structure for the SF_5^- anion in $CsSF_5$.

(ii) Identify the type of hybridization exhibited by sulfur in the SF_5^- anion.

(iii) Identify the geometry of the SF_5^- anion that is consistent with the Lewis structure drawn in part (b)(i).

(iv) Identify the oxidation number of sulfur in the compound $CsSF_5$.

2006 - #8

Suppose that a stable element with atomic number 119, symbol Q, has been discovered.

(a) Write the ground-state electron configuration for Q, showing only the valence-shell electrons.

(b) Would Q be a metal or a nonmetal? Explain in terms of electron configuration.

(c) On the basis of periodic trends, would Q have the largest atomic radius in its group or would it have the smallest? Explain in terms of electronic structure.

(d) What would be the most likely charge of the Q ion in stable ionic compounds?

(e) Write a balanced equation that would represent the reaction of Q with water.

(f) Assume that Q reacts to form a carbonate compound.

(i) Write the formula for the compound formed between Q and the carbonate ion, CO_3^{2-}.

(ii) Predict whether or not the compound would be soluble in water. Explain your reasoning.

2006B - #6

$GeCl_4$ $SeCl_4$ ICl_4^- ICl_4^+

The species represented above all have the same number of chlorine atoms attached to the central atom.

(a) Draw the Lewis structure (electron-dot diagram) of each of the four species. Show all valence electrons in your structures.

(b) On the basis of the Lewis structures drawn in part (a), answer the following questions about the particular species indicated.

(i) What is the Cl – Ge – Cl bond angle in $GeCl_4$?

(ii) Is $SeCl_4$ polar? Explain.

(iii) What is the hybridization of the I atom in ICl_4^-?

(iv) What is the geometric shape formed by the atoms in ICl_4^+?

2006B - #7

Account for each of the following observations in terms of atomic theory and/or quantum theory.

(a) Atomic size decreases from Na to Cl in the periodic table.

(b) Boron commonly forms molecules of the type BX_3. These molecules have a trigonal planar structure.

(c) The first ionization energy of K is less than that of Na.

(d) Each element displays a unique gas-phase emission spectrum.

2007 - #6a-d

Answer the following questions, which pertain to binary compounds.

(a) In the box provided below, draw a complete Lewis electron-dot diagram for the IF_3 molecule.

(b) On the basis of the Lewis electron-dot diagram that you drew in part (a), predict the molecular geometry of the IF_3 molecule.

(c) In the SO_2 molecule, both of the bonds between sulfur and oxygen have the same length. Explain this observation, supporting your explanation by drawing in the box below a Lewis electron-dot diagram (or diagrams) for the SO_2 molecule.

(d) On the basis of your Lewis electron-dot diagram(s) in part (c), identify the hybridization of the sulfur atom in the SO_2 molecule.

2007B - #2b-c

Answer the following problems about gases.

(b) A major line in the emission spectrum of neon corresponds to a frequency of 4.34×10^{14} s^{-1}. Calculate the wavelength, in nanometers, of light that corresponds to this line.

(c) In the upper atmosphere, ozone molecules decompose as they absorb ultraviolet (UV) radiation, as shown by the equation below. Ozone serves to block harmful ultraviolet radiation that comes from the Sun.

$$O_3(g) \xrightarrow{UV} O_2(g) + O(g)$$

A molecule of $O_3(g)$ absorbs a photon with a frequency of 1.00×10^{15} s^{-1}.

(i) How much energy, in joules, does the $O_3(g)$ molecule absorb per photon?

(ii) The minimum energy needed to break an oxygen-oxygen bond in ozone is 387 kJ mol^{-1}. Does a photon with a frequency of 1.00×10^{15} s^{-1} have enough energy to break this bond? Support your answer with a calculation.

2007B - #6

	First Ionization Energy (kJ mol^{-1})	Second Ionization Energy (kJ mol^{-1})	Third Ionization Energy (kJ mol^{-1})
Element 1	**1251**	**2300**	**3820**
Element 2	**496**	**4560**	**6910**
Element 3	**738**	**1450**	**7730**
Element 4	**1000**	**2250**	**3360**

The table above shows the first three ionization energies for atoms of four elements from the third period of the periodic table. The elements are numbered randomly. Use the information in the table to answer the following questions.

(a) Which element is most metallic in character? Explain your reasoning.

(b) Identify element 3. Explain your reasoning.

(c) Write the complete electron configuration for an atom of element 3.

(d) What is the expected oxidation state for the most common ion of element 2?

(e) What is the chemical symbol for element 2?

(f) A neutral atom of which of the four elements has the smallest radius?

2008 - #5

Using principles of atomic and molecular structure and the information in the table below, answer the following questions about atomic fluorine, oxygen, and xenon, as well as some of their compounds.

Atom	First Ionization Energy ($kJ\ mol^{-1}$)
F	1,681.0
O	1,313.9
Xe	?

(a) Write the equation for the ionization of atomic fluorine that requires 1,681.0 $kJ\ mol^{-1}$.

(b) Account for the fact that the first ionization energy of atomic fluorine is greater than that of atomic oxygen. (You must discuss both atoms in your response.)

(c) Predict whether the first ionization energy of atomic xenon is greater than, less than, or equal to the first ionization energy of atomic fluorine. Justify your prediction.

(d) Xenon can react with oxygen and fluorine to form compounds such as XeO_3 and XeF_4. In the boxes provided, draw the complete Lewis electron-dot diagram for each of the molecules represented below.

XeO_3 XeF_4

(e) On the basis of the Lewis electron-dot diagrams you drew for part (d), predict the following:

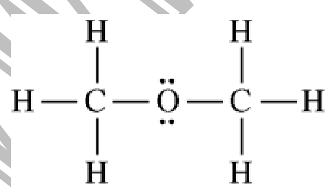
(i) The geometric shape of the XeO_3 molecule

(ii) The hybridization of the valence orbitals of xenon in XeF_4

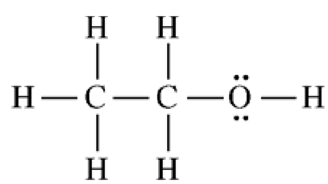
(f) Predict whether the XeO_3 molecule is polar or nonpolar. Justify your prediction.

2008 - #6 b-d

(b) Structures of the dimethyl ether molecule and the ethanol molecule are shown below. The normal boiling point of dimethyl ether is 250 K, whereas the normal boiling point of ethanol is 351 K. Account for the difference in boiling points. You must discuss both of the substances in your answer.



Dimethyl Ether



Ethanol

(c) SO_2 melts at 201 K, whereas SiO_2 melts at 1,883 K. Account for the difference in melting points. You must discuss both of the substances in your answer.

(d) The normal boiling point of Cl_2 (*l*) (238 K) is higher than the normal boiling point of HCl(*l*) (188 K). Account for the difference in normal boiling points based on the types of intermolecular forces in the substances. You must discuss both of the substances in your answer.

2009 - #3b & c

Initiating most reactions involving chlorine gas involves breaking the Cl–Cl bond, which has a bond energy of 242 kJ mol^{-1}.

(b) Calculate the amount of energy, in joules, needed to break a single Cl–Cl bond.

(c) Calculate the longest wavelength of light, in meters, that can supply the energy per photon necessary to break the Cl–Cl bond.

2009 - #5d

Reaction Y: $CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$ $\Delta H = +41$ kJ mol^{-1}

(d) For reaction Y at 298 K, which is larger: the total bond energy of the reactants or the total bond energy of the products? Explain.

2009 - #6

Answer the following questions related to sulfur and one of its compounds.

(a) Consider the two chemical species S and S^{2-}.

(i) Write the electron configuration (e.g., $1s^2\,2s^2$ . . .) of each species.

(ii) Explain why the radius of the S^{2-} ion is larger than the radius of the S atom.

(iii) Which of the two species would be attracted into a magnetic field? Explain.

(b) The S^{2-} ion is isoelectronic with the Ar atom. From which species, S^{2-} or Ar, is it easier to remove an electron? Explain.

(c) In the H_2S molecule, the H–S–H bond angle is close to 90°. On the basis of this information, which atomic orbitals of the S atom are involved in bonding with the H atoms?

(d) Two types of intermolecular forces present in liquid H_2S are London (dispersion) forces and dipole-dipole forces.

(i) Compare the strength of the London (dispersion) forces in liquid H_2S to the strength of the London (dispersion) forces in liquid H_2O. Explain.

(ii) Compare the strength of the dipole-dipole forces in liquid H_2S to the strength of the dipole-dipole forces in liquid H_2O. Explain.

2009B - 5a & c

Answer the following questions about nitrogen, hydrogen, and ammonia.

(a) Draw the complete Lewis electron-dot diagrams for N_2 and NH_3.

(c) Given that ΔH°_{298} for the reaction is −92.2 kJ mol^{-1}, which is larger, the total bond dissociation energy of the reactants or the total bond dissociation energy of the products? Explain.