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. **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$**

(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.

for the two electrons in the 4s: 4, 0, 0, -½ and 4, 0, 0, +½

for the three electrons in the 4p: 4, 1, -1, -½; 4, 1, -1, +½ and 4, 1, 0, -½

(c) Is an isolated arsenic atom in the ground state paramagnetic or diamagnetic? Explain briefly.

Paramagnetic. It has three unpaired electrons.

(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.

Na_3As - each Na gives up one electron to the As, the As has a complete octet and the sodium atoms are ionically bonded to the arsenic

$AsCl_3$ - the three chlorines each have one half-filled orbital and the arsenic has three. So three covalent bonds are created and the As has one non-bonding pair to make a pyramidal structure.

AsF_5 - fluorine is so electronegative that it draws the two electrons of the non-bonding pair of $AsCl_3$ into bonding. A 4d orbital is involved in the sp^3d hybridization, yielding a trigonal bipyramidal shape.

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.

Mg $1s^2\ 2s^2\ 2p^6\ 3s^2$

Ar $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

(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.

Valence electrons for Mg and Ar are in the same principal energy level, but Ar is the smaller and it has a greater nuclear charge. Therefore, the first and second ionization energies of Mg are less than those of Ar. The removal of the third electron from Mg involves the n = 2 energy level, and this electron experiences a very large nuclear charge.

(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). **$MgCl_2$ forms, since Mg readily loses its two valence electrons to form a stable configuration.**

(d) Element Q has the following first three ionization energies [in kJ/mole]:

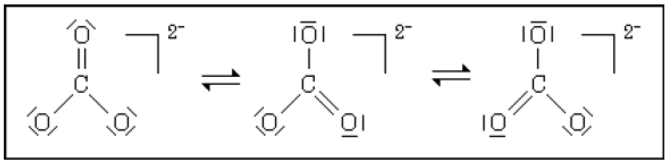
	I_1	I_2	I_3
Q	496	4568	6920

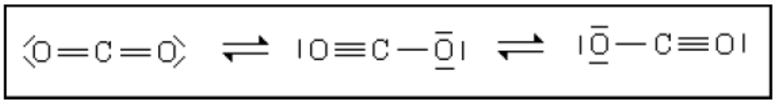
What is the formula for the most likely compound of element Q with chlorine? Explain the choice of formula on the basis of the ionization energies.

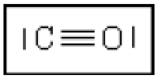
The formula is QCl. A very high second ionization energy indicates that there is only one valence electron.

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.

CO has the shortest bond distance because it has the greatest number of electrons between carbon and oxygen.

(c) Predict the molecular shapes for the three species. Explain how you arrived at your predictions.

CO_3^{2-} is a trigonal planar. There are 3 bonded pairs of electrons and no lone pairs around C. To minimize the repulsion of the bonded pairs there will be 120° bond angles with all the atoms in the same plane, or C is sp^2 hybridized with a delocalized pi bond, therefore, CO_3^{2-} must be planar.

CO_2 is linear. It has 2 bonded pairs and no lone pairs of electrons. Or C uses two sp hybridized orbitals.

CO is linear. Two points or atoms make a straight line.

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.

Physical properties	**Metals**	**Nonmetals**
Melting points	**Relatively high**	**Relatively low**
Electrical conductivity	**Good**	**Insulators**
Luster	**High**	**Little or none**
Physical state	**Most are solids**	**Gases, liquids, or solids**

Chemical properties	**Metals**	**Nonmetals**
Redox	**Reducing agents**	**Oxidizing or reducing agents**
Attraction to electrons	**Electropositive**	**Electronegative**
Acid/base character	**Oxides basic or amphoteric**	**Oxides acidic**
Reactivity	**React with nonmetals**	**React with both metals and nonmetals**

Electron configurations: Metals: Valence electrons in s or d sublevels of their atoms. (A few heavy elements have atoms with one or two electrons in p sublevels.) Nonmetals: Valence electrons in the s and p sublevels of their atoms. There are more metals than nonmetals because filling d orbitals in a given energy level involves the atoms of ten elements and filling the f orbitals involves the atoms of 14 elements. In the same energy levels, the maximum number of elements with atoms receiving p electrons is six.

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.

H_2 and C_3H_8 have low melting points because the forces involved are the weak van der Waals (or London) forces.

HF has a higher melting point because intermolecular hydrogen bonding is important.

CsI and LiF have still higher melting points because ionic lattice forces must be overcome to break up the crystals and the ionic forces are stronger than van der Waals forces and hydrogen bonds.

SiC is an example of a macromolecular substance where each atom is held to its neighbors by a very strong covalent bond.

(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.

C_3H_8 and H_2: There are more interactions per molecule in C_3H_8 than in H_2. Or C_3H_8 is weakly polar and H_2 is nonpolar.

LiF and CsI: The smaller ions in LiF result in a higher lattice energy than CsI has. Lattice energy U is proportional to $1 / (r^+ + r^-)$

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.

The radii of the alkali metal ions increase with increasing atomic number because:

(i) the principle quantum number (or shell or energy level) increases

(ii) there is an increase in shielding (or the number of orbitals 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)

The chloride ion is larger than the chlorine atom because:

(i) electron- electron repulsion increases (or shielding increases or the electron-proton ratio increases or the effective nuclear charge decreases)

(ii) an extra electron generally increases the size

(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)

The ionization energy of Mg is greater than that for Al because:

(i) the 3p orbital is at a higher energy than the 3s orbital (or the electron in Al is shielded from the nucleus more completely by the 3s electron than the 3s electrons shield one another from the nucleus)

(ii) a 3p electron is easier to remove than a 3s electron

(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)

The much greater difference between the 2nd and 3rd ionization energies in Mg (relative to the difference between the 1st and 2nd) is due to the 3rd electron being removed from the 2p subshell after the first 2 were removed from the 3s subshell.

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.

It is impossible to determine (or measure) both the position and momentum of any particle (or object, or body) simultaneously.

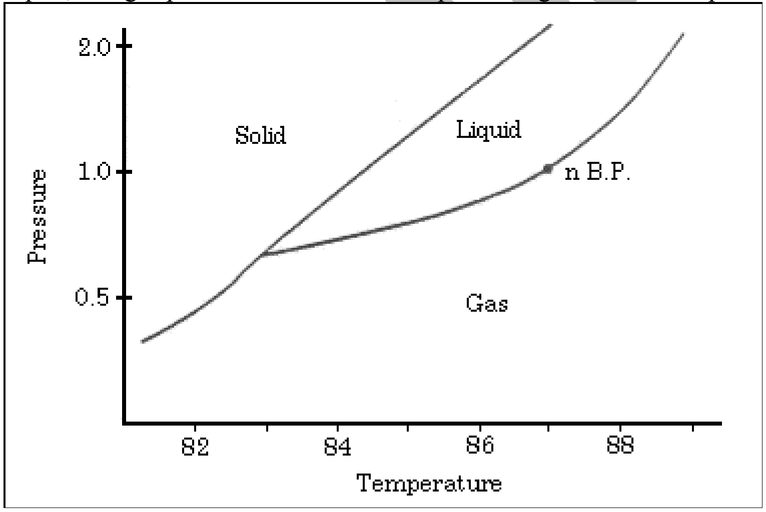
(b) What aspect of the Bohr Theory of the atom is considered unsatisfactory as a result of the Heisenberg uncertainty principle?

Bohr postulated that the electron in a H atom travels about the nucleus in a circular orbit and has a fixed angular momentum. This violates the Heisenberg uncertainty theory because Heisenberg viewed the electron as traveling as a wave function.

(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.
All matter exhibits both particulate and wave properties. Large matter exhibit predominately particulate properties. The associated wavelength is so small that it is not observed. Very small "bits of matter", such as photons, while showing some particulate properties exhibit predominately wave properties. Pieces of matter with intermediate mass, such as electrons, show clearly both the particulate and wave properties of matter.

1988 - #5 Average score: 2.99 out of 8
Using principles of chemical bonding and/or intermolecular forces, explain each of the following.
(a) Xenon has a higher boiling point than neon has.
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) Solid copper is an excellent conductor of electricity, but solid copper chloride is not.
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 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.
SiO_2 is a covalent network solid. There are strong bonds many of which must be broken simultaneously to volatize 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) Molecules of NF_3 are polar, but those of BF_3 are not.
A lone pair of electrons on the central atom results in a pyramidal shape. The dipoles don't cancel, thus the molecule is polar. 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.

1988 - #8 Average Score: 3.82 out of 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.
The argon sublimes.
(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.
The argon vaporizes.

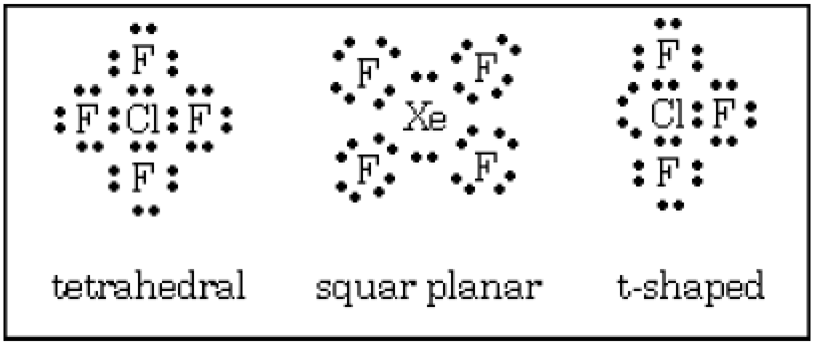
(d) Does the liquid phase of argon have a density greater than, equal to, or less than the density of the solid phase? Explain your answer, using information given in the introduction to this question.

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.

1989 - #5 Average Score: 2.7 out of 8

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.

CF_4 - 4 bonding pairs around the 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 atoms.

ClF_3 - 3 bonding pairs and 2 lone pairs give trigonal bipyramid with lone pairs in equatorial positions 120° apart.

1989 - #6 Average Score: 1.6 out of 8

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.

Alkali metals have metallic bonds: cations in a sea of electrons. As cations increase in size (Li to Cs), charge density decreases and attractive forces (and melting points) decreases.

(b) Using bonding principles, account for the increase in the melting points of the halogens.

Halogen molecules are held in place by dispersion (van der Waals) forces: bonds due to temporary dipoles caused by polarization of electron clouds. As molecules increase in size (F_2 to I_2), the larger electrons clouds are more readily polarized, and the attractive forces (and melting points) increase.

(c) What is the expected trend in the melting points of the compounds LiF, NaCl, KBr, and CsI? Explain this trend using bonding principles.

Melting point order: LiF > NaCl > KBr > CsI
Compounds are ionic
Larger radii of ions as listed
Larger radii leads to smaller attraction and lower melting points.

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.

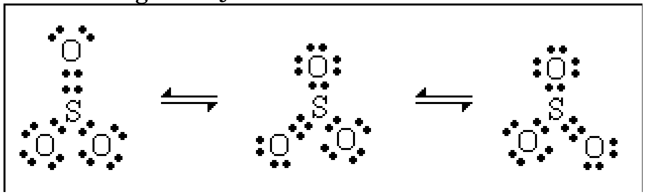
C_2H_4 has a multiple bond; C_2H_6 has a single bond. Multiple bonds are stronger and therefore shorter than single bonds.

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

NH_3 has 3 bonding pairs and 1 lone pair of electrons.
Bond pairs are forced together because the repulsion between the lone pair and the bond pairs is greater than that between bond pairs.

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

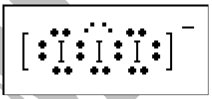
The bonding in SO_3 can be described as a combination of 3 resonance forms of 1 double and single bonds.



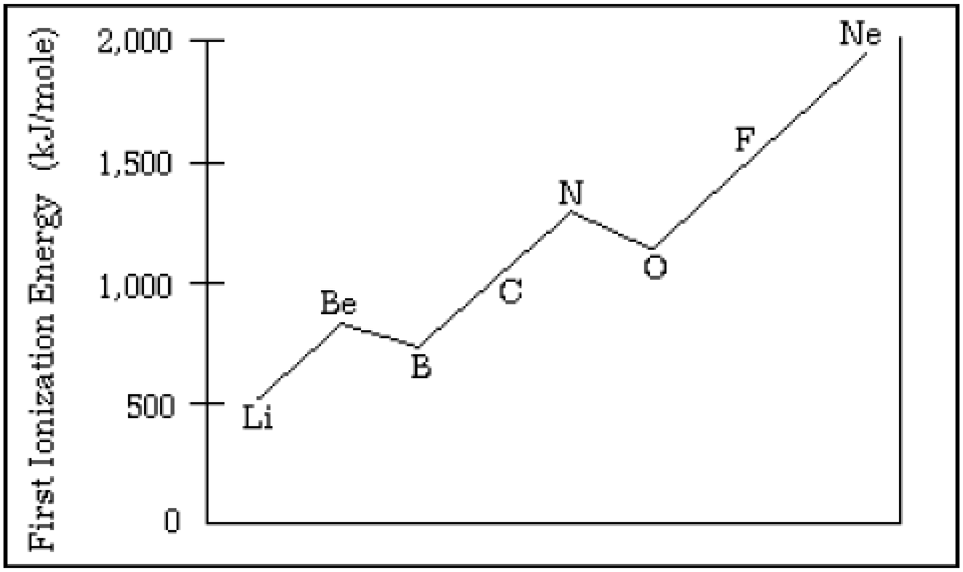
The actual structure is intermediate between the 3 resonance forms, having 3 bonds which are equal and stronger (therefore shorter) than a S-O single bond.

(d) The I_3^- ion is linear.

The central I atom has 3 lone pairs and 2 bond pairs around it. To minimize repulsion, the 3 lone pairs are arranged in a trigonal plane at right angles to the I-I-I axis.



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.

Across the period from Li to Ne the number of protons is increasing in the nucleus hence the nuclear charge is increasing with a consequently stronger attraction for electrons and an increase in I.E.

(b) The first ionization energy of B is lower than that of Be.

The electron ionized in the case of Be is a 2s electron whereas in the case of B it is a 2p electron. 2p electrons are higher in energy than 2s electrons because 2p electrons penetrate the core to a lesser degree.

(c) The first ionization energy of O is lower than that of N.

The electron ionized in O is paired with another electron in the same orbital, whereas in N the electron comes from a singly-occupied orbital. The ionization energy of the O electron is less because of the repulsion between two electrons in the same orbital.

(d) Predict how the first ionization energy of Na compares to those of Li and of Ne. Explain.

The ionization energy of Na will be less than those of both Li and Ne because the electron removed comes from an orbital which is farther from the nucleus, therefore less tightly held.

1992 - #8 Average Score: 3.0 out of 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.

K conducts because of its metallic bonding or "sea" of mobile e's (or "free" e's) KNO_3 does not conduct because it is ionically bonded and has immobile ions (or imm. e's)

(b) $SbCl_3$ has a measurable dipole moment, whereas $SbCl_5$ does not.

$SbCl_3$ has a measurable dipole moment because it has a lone pair of e's which causes a dipole or its dipoles do not cancel or it has a trigonal pyramidal structure or clear diagram illustrating any of the above $SbCl_5$ has no dipole moment because its dipoles cancel or it has a trigonal bipyramidal structure

(c) The normal boiling point of CCl_4 is 77 °C, whereas that of CBr_4 is 190 °C.

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

(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.

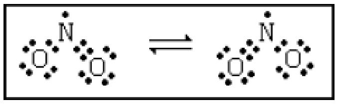
NaI has greater aqueous solubility than I_2 because NaI is ionic (or polar) whereas I_2 is nonpolar (or covalent). H_2O, being polar, interacts with the ions of NaI but not with I_2. (Like dissolves like accepted if polarity of H_2O clearly indicated.)

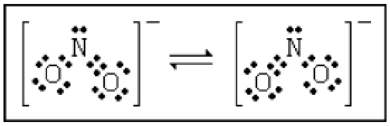
1992 - #9 Average Score: 2.5 out of 8

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.

$NO_2^- < NO_2 < NO_2^+$

(c) Select one of the species and give the hybridization of the nitrogen atom in it.

NO_2^+ is sp
NO_2 is sp^2
NO_2^- is sp^2

(d) Identify the only one of the species that dimerizes and explain what causes it to do so.

NO_2 will dimerize because it contains an odd electron that will pair readily with another, giving N_2O_4.

1993 - #6 Average Score: 2.1 out of 8

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.

Electron configuration of Na and Mg
Octet / Noble gas stability comparison of Na and Mg
Energy difference explanation between Na and Mg
Size difference explanation between Na and Mg
Shielding/effective nuclear charge discussion

(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.

Correct direction and explanation of the following:
shielding differences
energy differences
# of proton/ # of electron differences

(c) A sample of solid nickel chloride is attracted into a magnetic field, whereas a sample of solid zinc chloride is not.

(i) Ni unpaired electrons. paramagnetic
(ii) Zn paired electrons/ diamagnetic
(iii) Ni unpaired electrons/ Zn paired electrons
(iv) Ni paramagnetic/ Zn diamagnetic

Explanation must include Orbital discussion and Hund's Rule

(d) Phosphorus forms the fluorides PF_3 and PF_5, whereas nitrogen forms only NF_3.

Expanded octet or sp^3d hybrid of phosphorous

Lack of d orbitals in nitrogen

Nitrogen is too small to accommodate (or bond) 5 Fluorines or 5 bonding sites

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.

Ca^{2+} has fewer electrons, thus it is smaller than Ca. The outermost electron in Ca is in a 4s orbital, whereas the outermost electron in Ca^{2+} is in a 3p orbital.

(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.

U for CaO is more negative than U for K_2O, so it is more difficult to break up the CaO lattices (stronger bonds in CaO). Ca^{2+} is smaller than K^+, so internuclear separations (between cations and O_2^-) are less. Ca^{2+} is more highly charged than K^+, thus cation-O_2 bonds are stronger

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.

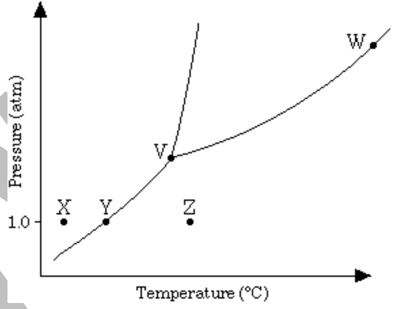
  Ca has ore protons and is smaller. The outermost electrons are more strongly held by the nuclear charge of Ca.
- (ii) their second ionization energies.

  The outermost electrons in Ca are in the 4s, which is a higher energy orbital (more shielded) than the 3p electrons in K.

(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.

The highest energy (outermost) electrons in Al is in a 3p orbital, whereas that electron in Mg is in a 3s orbital. The 3p electron in Al is of higher energy (is more shielded) than is the 3s electron in Mg.

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?

V is the triple point (or point where 3 phases coexist). Solid, liquid, and vapor (or 3 phases) are in equilibrium.

(b) What does each point on the curve between V and W represent?

Each point on the curve represents the temperature and pressure where the liquid and vapor (or 2 phases) coexist. At these temperatures and pressures, the two phases are in equilibrium. The points represent the vapor pressure of the liquid as a function of temperature. The points represent the boiling points of the liquid as a function of (applied) pressure.

(c) Describe the changes that the system undergoes as the temperature slowly increases from X to Y to Z at 1.0 atmosphere.

Changes: sublimation, phase change, energy change, density change, and entropy change

Point Y: change in phase occurs specifically at Y

(d) In a solid-liquid mixture of this substance, will the solid float or sink? Explain.

The solid will sink. The positive slope of the solid/liquid equilibrium curve indicates that the solid is denser than the liquid.

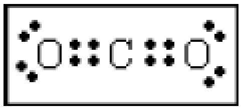
1995 - #7

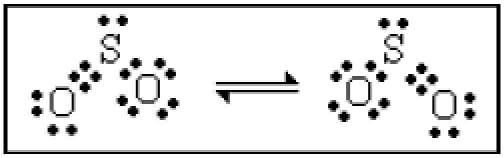
Explain the following in terms of the electronic structure and bonding of the compounds considered.

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

Oxygen has unpaired electrons and is paramagnetic; thus it is attracted to a magnetic field.

(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.





The CO_2 molecule is linear, thus bond dipoles cancel; the SO_2 molecule is bent, resulting in a net dipole.

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

Co^{2+} has a partially filled d subshell, whereas the d subshell in Zn^{2+} is filled. As a result, an electron can be excited between d orbitals in Co^{2+}, causing visible light absorption; this cannot happen in Zn^{2+}.

(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.

Si has all of its valence electrons localized in covalent bonds (in a network lattice), hence it is a poor conductor. Introduction of As atoms with their "extra" electrons into the Si lattice allows for an increase in conductivity.

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.

Hydrogen bonding (or dipole-dipole attraction) in HF is greater than it is in HCl

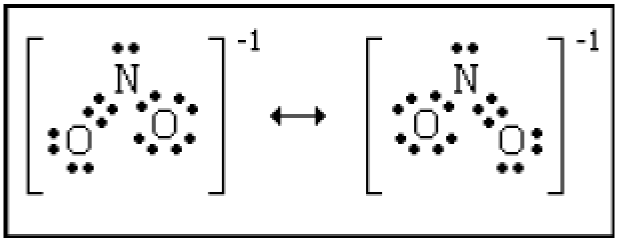
(b) Molecules of AsF_3 are polar, whereas molecules of AsF_5 are nonpolar.

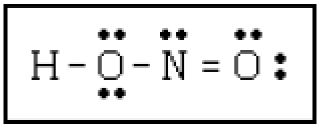
AsF_3 has a trigonal pyramid shape and bond dipoles do NOT cancel (or asymmetric molecule)

AsF_5 has a trigonal bipyramid shape and bond dipoles cancel (or symmetric shape)

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

NO_2^- has resonance structures. HNO_2 has no resonance structures.





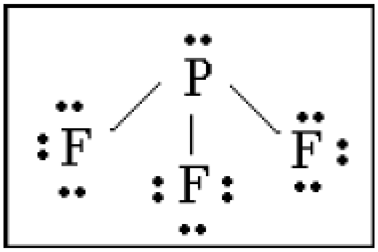
(d) For sulfur, the fluorides SF_2, SF_4, and SF_6 are known to exist, whereas for oxygen only OF_2 is known to exist.

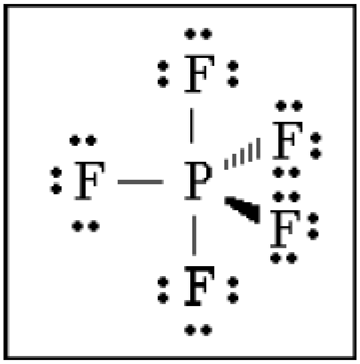
Sulfur uses d orbitals (or expanded octet), oxygen has no d orbitals in its valence shell. Sulfur is a larger atom, can accommodate more bonds.

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.





PF_3 is trigonal pyramidal. PF_5 is trigonal bipyramidal.

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

The PF_3 molecule is polar. The three P-F dipoles do not cancel, or, the lone pair on P leads to asymmetrical distribution of charge.

(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

NF_5 does not exist because no 2d orbitals exist for use in bonding. N is too small to accommodate 5 bonding pairs. AsF_5 does exist because 4d orbitals are available for use in bonding. As can accommodate an expanded octet using d orbitals

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.

Response must contain a cogent discussion of the forces between the nucleus and the outermost (or "ionized") electron. For example, a discussion of "the outermost electron on K..." should include one of the following:

i. it is farther from nucleus than the outermost electron on Li
ii. it is more shielded from the nucleus (or "experiences a lower effective nuclear charge") than the outermost electron on Li
iii. it is in a higher energy orbital (4s) than tne outermost electron on Li (2s)."

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

Nitrogen has one less proton than oxygen. Nitride and oxide ions are isoelectronic. In nitride ion the electron/proton ratio is greater, causing more repulsion; thus, nitride is the larger ion.

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

A Zn atom has more protons (10 more) than an atom of Ca. Electrons in d orbitals of Zn have a lower principal quantum number; thus, they are not in orbitals that are farther from the nucleus.

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

Correct identification of the orbitals involved (2s versus 2p). Clear statement that the two orbitals have different energies

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. **6.06×10^{14} sec^{-1}**
(ii) Calculate the energy, in J, of a photon of the light. **4.02×10^{-19} J**
(iii) Calculate the minimum energy, in kJ mol^{-1}, of the Cl-Cl bond. **242 kJ/mol**

(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.

Energy is emitted. The $n = 6$ state is at a higher energy than the $n = 2$ state. Going from a high energy state to a low energy state means that energy must be emitted.

(ii) Calculate the wavelength, in nm, of the radiation associated with the spectral line. **411 nm**

(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.

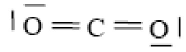
The positive charge holding the electron is greater for He^+, which has a 2+ nucleus, than for H with its 1+ nucleus. The stronger attraction means that it requires more energy for the electron to move to higher energy levels. Therefore, transitions from high energy states to lower states will be more energetic for He^+ than for H.

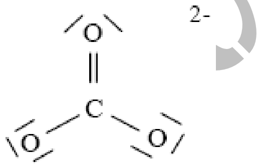
1999 - #8

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-}.

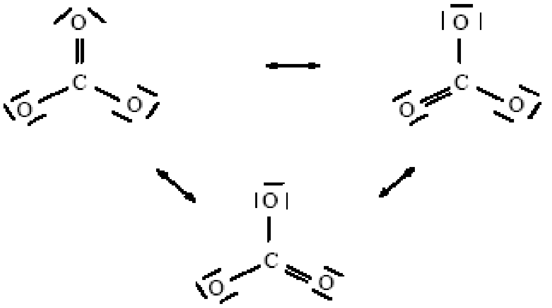
a. Draw the complete Lewis electron-dot structure for each species.





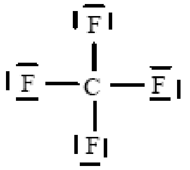
b. Account for the fact that the carbon-oxygen bond length in CO_3^{2-} is greater than the carbon-oxygen bond length in CO_2.

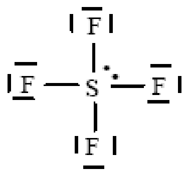
In CO_2, the C−O interactions are double bonds. In CO_3^{2-} the C−O interactions are resonance forms (or figures to the right.) The carbon-oxygen bond length is greater in the resonance forms than in the double bonds.



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.

CF_4 has a tetrahedral geometry, so the bond dipoles cancel, leading to a nonpolar molecule. With five pairs of electrons around the central S atom, SF_4 exhibits a trigonal bipyramidal electronic geometry, with the lone pair of electrons. In this configuration, the bond dipoles do not cancel, and the 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.

The isotopes have the same number (34) of protons, but a different number of neutrons.

(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.

$\boldsymbol{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4}$

Since there are three different 4p orbitals, there must be two unpaired electrons. There must be some explanation of Hund's rule, and a orbital diagram.

(c) In terms of atomic structure, explain why the first ionization energy of selenium is

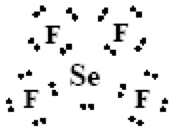
(i) less than that of bromine (atomic number 35), and

The ionized electrons in both Se and Br are in the same energy level, but Br has more protons than Se, so the attraction to the nucleus is greater.

(ii) greater than that of tellurium (atomic number 52).

The electron removed from a Te atom is in a 5*p* orbital, while the electron removed from an Se atom is in a 4*p* orbital. The 5*p* orbital is at a higher energy than the 4*p* orbital, thus the removal of an electron in a 5*p* orbital requires less energy.

(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.



The SeF_4 molecule is polar, because the polarities induced by the bonds and the lone pair of electrons do not cancel.

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).

NH_3 has hydrogen bonding between molecules (*or* dipole-dipole interactions between molecules), and CH_4 has London dispersion forces. The intermolecular forces in NH_3 are stronger than those in CH_4.

(b) At 25°C and 1.0 atm, ethane (C_2H_6) is a gas and hexane (C_6H_{14}) is a liquid.

C_2H_6 and C_6H_{14} both have London dispersion forces. The forces in C_6H_{14} are stronger because the molecule is larger and more polarizable.

(c) Si melts at a much higher temperature (1,410°C) than Cl_2 (-101°C).

Si is a network covalent solid (*or* a macromolecule) with strong covalent bonds between atoms. Cl_2 has discrete molecules with weak London dispersion forces between the molecules. Therefore, more energy is required to break the stronger bonds of Si than the weak intermolecular forces of Cl_2. (d) MgO melts at a much higher temperature (2,852°C) than NaF (993°C).

MgO and NaF are both ionic solids (*or* ions are listed to indicate this).

+2 and -2 charges in MgO result in a greater attraction between ions than the +1 and -1 charges in NaF (*or* according to Coulomb's Law, the attraction between +2 ions and -2 ions is greater than that between +1 ions and -1 ions,

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.

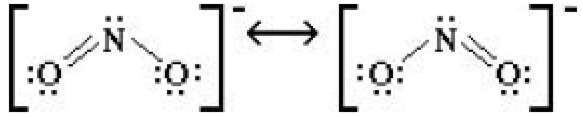
(a) The atomic radius of Li is larger than that of Be.

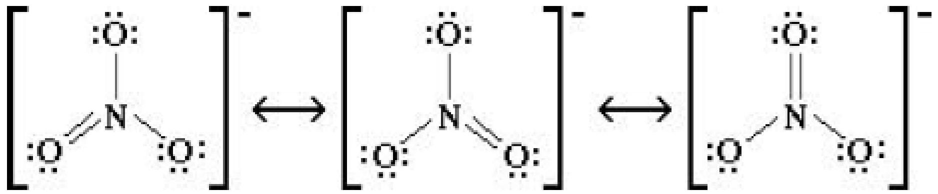
Both Li and Be have their outer electrons in the same shell (and/or they have the same number of inner core electrons shielding the valence electrons from the nucleus). However, Be has four protons and Li has only three protons. Therefore, the effective nuclear charge experienced (attraction experienced) by the valence (outer) electrons is greater in Be than in Li, so Be has a smaller atomic radius.

(b) The second ionization energy of K is greater than the second ionization energy of Ca.
The second electron removed from a potassium atom comes from the third level (inner core). The second electron removed from a calcium atom comes from the fourth level (valence level). The electrons in the third level are closer to the nucleus so the attraction is much greater than for electrons in the fourth level.
(c) The carbon-to-carbon bond energy in C_2H_4 is greater than it is in C_2H_6.
C_2H_4 has a double bond between the two carbon atoms, whereas C_2H_6 has a carbon-carbon single bond. More energy is required to break a double bond in C_2H_4 than to break a single bond in C_2H_6; therefore, the carbon-to-carbon bond energy in C_2H_4 is greater.
(d) The boiling point of Cl_2 is lower than the boiling point of Br_2.
Both Cl_2 and Br_2 are nonpolar, and the only intermolecular attractive forces are London dispersion forces. Since Br_2 has more electrons than Cl_2, the valence electrons in Br_2 are more polarizable. The more polarizable the valence electrons, the greater the dispersion forces and the higher the boiling point.

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^-.





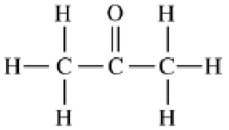
According to the Lewis electron-dot diagram, two resonance structures are required to represent the bonding in the NO_2^- ion. The effective number of bonds between N and O is 1.5.
Three resonance structures are required to represent the bonding in the NO_3^- ion. The effective number of bonds between N and O is 1.33. The greater the effective number of bonds, the shorter the N–O bond length.
(b) The CH_2F_2 molecule is polar, whereas the CF_4 molecule is not.
The molecular geometry in both CH_2F_2 and CF_4 is tetrahedral (or the same). The C-F bond is polar. In CF_4, the molecular geometry arranges the C-F dipoles so that they cancel out and the molecule is nonpolar. The C-H bond is less polar than the C-F bond. The two C-H dipoles do not cancel the two C-F dipoles in CH_2F_2.
(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.
The carbon atoms in C_2H_4 have a molecular geometry around each carbon atom that is trigonal planar (AX_3), so all six atoms are in the same plane. The carbon atoms in C_2H_6 have a molecular geometry that is tetrahedral (AX_4), so the atoms are not all in the same plane. OR The carbon-carbon double bond in C_2H_4 results in a planar molecule whereas the carbon-carbon single bond in C_2H_6 results in a non-planar (tetrahedral) site at each carbon atom.
(d) The shape of a PF_5 molecule differs from that of an IF_5 molecule.
In PF_5, the molecular geometry is trigonal bipyramidal because the phosphorus atom has five bonding pairs of electrons and no lone pairs of electrons. IF_5 has square pyramidal molecular geometry. The central iodine atom has five bonding pairs of electrons and one lone pair of electrons. The presence of the additional lone pair of electrons on the central iodine atom means the molecular geometry is different.

(e) $HClO_3$ is a stronger acid than HClO.
According to the formula for HOCl and $HOClO_2$, there are two additional terminal, electronegative oxygen atoms attached to the central chlorine atom. These additional terminal oxygen atom stabilize the negative charge on the anion ClO_3^- compared to ClO^-. The result is to reduce the electrostatic attraction between the H^+ and ClO_x^-. OR The two additional terminal electronegative O atoms bonded to the chlorine atom of ClO_3^- pull electron density away from the central chlorine atom. The net result is to weaken the H-O bond. Since HOCl has no additional terminal O atoms, its H-O bond is stronger. The weaker the H-O bond, the stronger the acid.

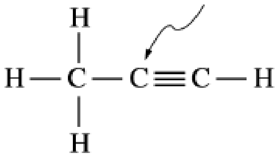
2003 - #8 a & d

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.
The C – C –C bond angle is 120°
(d) Given the structural formula for propyne below,



(i) indicate the hybridization of the carbon atom indicated by the arrow in the structure above;
sp hybridization
(ii) indicate the total number of sigma (σ) bonds and the total number of pi (π) bonds in the molecule.
6 sigma bonds and 2 pi bonds

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.
Both Ca^{2+} and Cl^- ions have 18 electrons. Their electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. However, they differ by the number of protons in the nucleus. Calcium has 20 protons and chlorine has 17 protons.
The valence electrons are shielded by the same number of electrons in each ion (10), so the effective nuclear charge (ENC) experienced by the valence electrons in Ca^{2+} is +10 and for Cl^- the ENC is +7. The valence electrons in Cl^- experience a smaller attraction to the nucleus due to the smaller nuclear charge, so Cl^- has the larger ionic radius.
(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.
Binary compounds of carbon exhibit covalent character (property of a nonmetallic element), whereas binary compounds of lead exhibit ionic character (property of a metallic element). Oxides of carbon, when dissolved in water, are acidic (property of a nonmetallic element), whereas oxides of lead, when added to water, are basic (property of a metallic element). Carbon is a poor thermal conductor (property of a nonmetallic element), whereas lead is a very good thermal conductor (property of a metallic element).

(c) Compounds containing Kr have been synthesized, but there are no known compounds that contain He.
Helium has a filled shell (the first shell), so does not tend to lose or gain electrons. Therefore, helium does not react. Krypton, while having filled 4s and 4p sublevels, has empty 4d and 4f sublevels. These empty orbitals affect the reactivity of Kr. Note: Also acceptable is a comparison of the ionization energies of helium, and krypton and then the justification for krypton being more reactive.
(d) The first ionization energy of Be is 900 kJ mol^{-1}, but the first ionization energy of B is 800 kJ mol^{-1}.
The electron configuration for Be is $1s^2\,2s^2$, whereas the electron configuration for B is $1s^2\,2s^2\,2p^1$. The first electron removed in boron is in a 2p subshell, which is higher in energy than the 2s subshell, from which the first electron is removed in beryllium. The higher in energy the subshell containing the electron to be removed (ionized), the lower the ionization energy.

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.
Both F_2 and I_2 are nonpolar, so the only intermolecular attractive forces are London dispersion forces. I_2 is solid because the electrons in the I_2 molecule occupy a larger volume and are more polarizable compared to the electrons in the F_2 molecule. As a result, the dispersion forces are considerably stronger in I_2 compared to F_2.
(b) The melting point of NaF is 993°C, whereas the melting point of CsCl is 645°C.
Both NaF and CsCl are ionic compounds with the same charges on the cations and anions. The ionic radius of Na^+ is smaller than the ionic radius of Cs^+ and the ionic radius of F^- is smaller than the ionic radius of Cl^-. Therefore, the ionic centers are closer in NaF than in CsCl. Melting occurs when the attraction between the cation and the anion are overcome due to thermal motion. Since the lattice energy is inversely proportional to the distance between the ion centers (Coulomb's Law), the compound with the smaller ions will have the stronger attractions and the higher melting point.
(c) The shape of the ICl_4^- ion is square planar, whereas the shape of the BF_4^- ion is tetrahedral.
The central iodine atom in ICl_4^- has four bonding pairs and two lone pairs of electrons on the central iodine atom, so the molecular geometry is square planar. BF_4^- has four bonding pairs and no lone pairs on the central boron atom, so the molecular geometry is tetrahedral.
(d) Ammonia, NH_3, is very soluble in water, whereas phosphine, PH_3, is only moderately soluble in water.
Ammonia has hydrogen-bonding intermolecular forces, whereas phosphine has dipole-dipole and/or dispersion intermolecular forces. Water also has hydrogen-bonding intermolecular attractive forces. Ammonia is more soluble in water than phosphine because ammonia molecules can hydrogen-bond with water molecules, whereas phosphine molecules cannot hydrogen-bond with water molecules.

2004 - #8 a & b

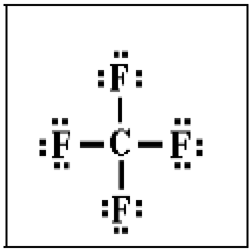
Answer the following questions about carbon monoxide, CO(*g*), and carbon dioxide, CO_2(*g*). Assume that both gases exhibit ideal behavior
(a) Draw the complete Lewis structure (electron-dot diagram) for the CO molecule and for the CO_2 molecule.

(b) Identify the shape of the CO_2 molecule.
CO_2 is linear.

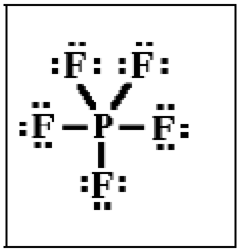
2005 - #6

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.

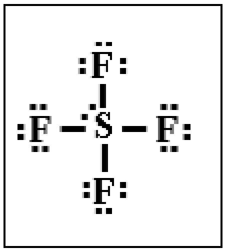
CF_4



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?

109.5°

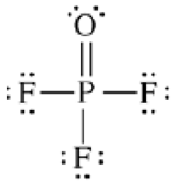
(ii) What is the hybridization of the valence orbitals of P in PF_5?

dsp^3

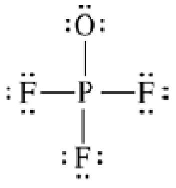
(iii) What is the geometric shape formed by the atoms in SF_4?

seesaw

(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?

4 sigma bonds and 1 pi bond

(ii) Which one of the two structures best represents a molecule of OPF_3? Justify your answer in terms of formal charge.

Structure 1 is the better structure because all of its atoms have a formal charge of zero.

P: 5 – 5 – 0 = 0
F: 7 – 1 – 6 = 0
O: 6 – 2 – 4 = 0

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.

NH_3 has dispersion forces and hydrogen-bonding forces. NF_3 has dispersion forces and dipole-dipole forces.

(ii) Account for the difference in the boiling points of the substances.

The higher boiling point for NH_3 is due to the greater strength of the hydrogen-bonding intermolecular attractive forces among NH_3 molecules compared to that of the dipole-dipole attractive forces among NF_3 molecules.

(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.

Both KCl and NaCl have ionic bonds.

(ii) Account for the difference in the melting points of the substances.

The difference in the melting points is due to the different strengths of ionic bonding in the substances. The charges on the cations and anions are the same in both compounds, therefore the relative size of the ions is the determining factor. Since Na^+ has a smaller ionic radius than K^+, the lattice energy of NaCl is higher than that of KCl. Thus more energy is required to overcome the ionic forces in solid NaCl than in solid KCl, and NaCl has the higher melting point.

(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.

The valence electron is located in the $n = 3$ level for all three atoms.

(ii) Explain the reasons for the trend in first ionization energies.

Because the valence electrons in all three elements are shielded by the same number of inner core electrons and the nuclear charge increases going from Si to P to Cl, the valence electrons feel an increasing attraction to the nucleus going from Si to P to Cl. Valence electrons having a greater attraction to the nucleus, as in Cl , will be more difficult to remove, so Cl has the highest ionization energy. P has the second highest ionization energy, and Si has the lowest ionization energy.

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.

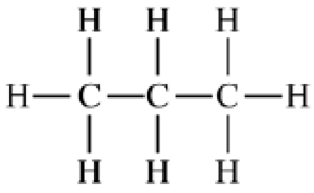
$:\ddot{S}=C=\ddot{S}:$

(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.

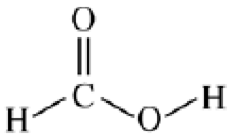
The carbon-to-selenium bond length in CSe_2 is greater than the carbon-to-sulfur bond length in CS_2. Because the valence electrons in Se are in a higher shell ($n = 4$) than the valence electrons in S ($n = 3$), Se has a larger atomic radius than S has. Therefore, the carbon to selenium bond length is greater than the carbon to sulfur bond length.

(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.

The carbon-to-selenium bond energy in CSe_2 is less than the carbon-to-sulfur bond energy in CS_2 because Se has a larger atomic radius than S. Because Se is a larger atom, the orbital overlap between the Se and C will be smaller than the orbital overlap between S and C.



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	**Propane has dispersion forces.**
Methanoic acid	374 K	**Methanoic acid has dispersion forces and hydrogen bonding forces.**

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

Hydrogen bonding IMFs among methanoic acid molecules are much stronger than dispersion forces among propane molecules. The stronger the IMFs, the more energy it takes to overcome them. Therefore, 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.

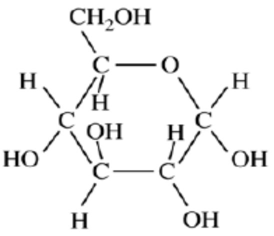
(a) The structures for glucose, $C_6H_{12}O_6$, and cyclohexane, C_6H_{12} , are shown below. Identify the type(s) of intermolecular attractive forces in

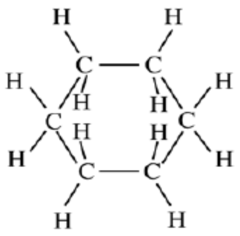
(i) pure glucose

Hydrogen bonding, dipole interactions and London dispersion forces.

(ii) pure cyclohexane

London dispersion forces





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

The hydroxyl groups in glucose molecules can form strong hydrogen bonds with the solvent (water) molecules, so glucose is soluble in water. In contrast, cyclohexane is not capable of forming strong intermolecular attractions with water (no hydrogen bonding), so the water cyclohexane interactions are not as energetically favorable as the interactions that already exist among polar water molecules.

(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.

In process 1, hydrogen bonds in liquid water are overcome to produce distinct water molecules in the vapor phase. In process 2, covalent bonds (sigma bonds) within water molecules must be broken to allow the oxygen atoms to recombine into molecular hydrogen and oxygen.

(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.

I disagree with the statement. Boiling is simply process 1, in which intermolecular forces are broken and the water molecules stay intact. No intramolecular forces or covalent bonds break in this process.

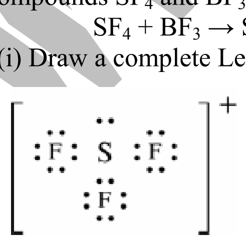
2006 - #7

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.

sp^3

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

trigonal pyramidal

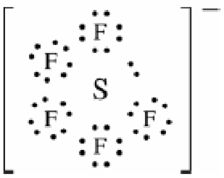
(iv) Predict whether the F–S–F bond angle in the SF_3^+ cation is larger than, equal to, or smaller than 109.5°. Justify your answer.

The F–S –F bond angle in the SF_3^+ cation is expected to be slightly smaller than 109.5° because the repulsion between the nonbonding pair of electrons and the S –F bonding pairs of electrons "squeezes" the F–S –F bond angles together slightly.

(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.

d^2sp^3

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

Square pyramidal

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

+4

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. **$8s^1$**

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

It would be a metal (OR an alkali metal). The valence electron would be held only loosely.

(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.

It would have the largest atomic radius in its group because its valence electron is in a higher principal shell.

(d) What would be the most likely charge of the Q ion in stable ionic compounds? **+1**

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

$2\ Q(s) + 2\ H_2O(l) \rightarrow 2\ Q^+(aq) + 2OH^-(aq) + H_2(g)$

(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-}.

Q_2CO_3

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

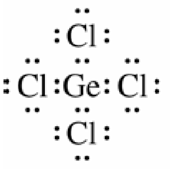
It would be soluble in water because all alkali metal carbonates are soluble.

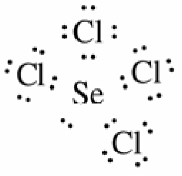
2006B - #6

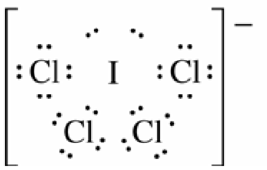
$GeCl_4$ $SeCl_4$ ICl_4^- ICl_4^+

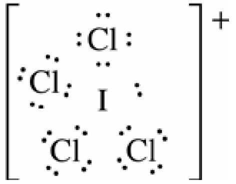
The species represented above all have the same number of chlorine atoms attached to the central atom.

(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 $GeCl_4$?

109.5°

(ii) Is $SeCl_4$ polar? Explain.

Yes. The $SeCl_4$ molecule is polar because the lone pair of nonbonding electrons in the valence shell of the selenium atom interacts with the bonding pairs of electrons, causing a spatial asymmetry of the dipole moments of the polar Se-Cl bonds. The result is a $SeCl_4$ molecule with a net dipole moment.

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

d^2sp^3

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

See-saw

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.

Across the periodic table from Na to Cl, the number of electrons in the s- and p- orbitals of the valence shell increases, as does the number of protons in the nucleus. The added electrons only partially shield the added protons, resulting in an increased effective nuclear charge. This results in a greater attraction for the electrons, drawing them closer to the nucleus, making the atom smaller.

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

Boron has three valence electrons, each of which can form a single covalent bond with X. The three single covalent bonds of the boron atom orient to minimize electron-pair interaction, resulting in bond angles of 120° and a trigonal planar structure.

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

Both Na and K have an s^1 valence-shell electron configuration (Na: [Ne] $3s^1$; K: [Ar] $4s^1$). The K atom valence electron has a higher *n* quantum number, placing it farther from the nucleus than the Na atom valence electron. The greater distance results in less attraction to the nucleus. Because its valence electron is less attracted to its nucleus, the K atom has the lower ionization energy.

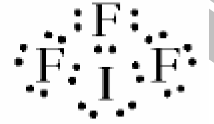
(d) Each element displays a unique gas-phase emission spectrum.

Each element has a unique set of quantized energy states for its electrons (because of its unique nuclear charge and unique electron configuration). As the electrons of an element absorb quanta of energy, they change to higher energy states (are excited) – during de-excitation, energy is released as EM radiation as the electrons cascade to lower energy states. Each photon of the EM radiation is associated with a specific wavelength ($\lambda = hc/E$), a flux of which produces the lines of the emission spectrum.

2007 - #6a-d

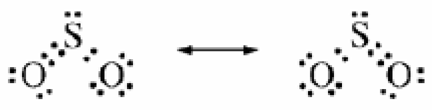
Answer the following questions, which pertain to binary compounds.

(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. **T-shaped**

(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. **sp^2**

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 x 10^{14} s^{-1}. Calculate the wavelength, in nanometers, of light that corresponds to this line. **690 nm**

(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.

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

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

(i) How much energy, in joules, does the $O_3(g)$ molecule absorb per photon? **6.63 ×10^{-19} J 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 x 10^{15} s^{-1} have enough energy to break this bond? Support your answer with a calculation.

$$\frac{6.63 \times 10^{-19}\ \text{J}}{1\ \text{photon}} \times \frac{6.022 \times 10^{23}\ \text{photons}}{1\ \text{mol}} \times \frac{1\ \text{kJ}}{10^3\ \text{J}} = 399\ \text{kJ mol}^{-1}$$

399 kJ mol^{-1} > 387 kJ mol^{-1}, therefore the bond can be broken.

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.

Element 2. It has the lowest first-ionization energy. Metallic elements lose electron(s) when they become ions, and element 2 requires the least amount of energy to remove an electron.

(b) Identify element 3. Explain your reasoning.

Magnesium. Element 3 has low first and second ionization energies relative to the third ionization energy, indicating that the element has two valence electrons, which is true for magnesium. (The third ionization of element 3 is dramatically higher, indicating the removal of an electron from a noble gas core.)

(c) Write the complete electron configuration for an atom of element 3. **$1s^2\ 2s^2\ 2p^6\ 3s^2$**

(d) What is the expected oxidation state for the most common ion of element 2? **1+**

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

(f) A neutral atom of which of the four elements has the smallest radius? **Element 1**

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	?

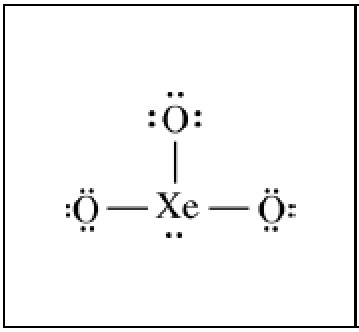
(a) Write the equation for the ionization of atomic fluorine that requires 1,681.0 kJ mol^{-1}. **$F(g) \rightarrow F^+(g) + e^-$**

(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.) **In both cases the electron removed is from the same energy level (2*p*), but fluorine has a greater effective nuclear charge due to one more proton in its nucleus (the electrons are held more tightly and thus take more energy to remove).**

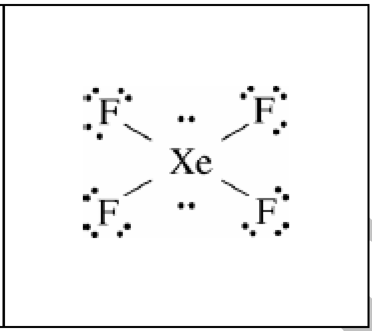
(c) Predict whether the first ionization energy of atomic xenon is greater than, less than, or equal to the first ionization energy of atomic fluorine. Justify your prediction. **The first ionization energy of Xe should be less than the first ionization energy of F. To ionize the F atom, an electron is removed from a 2*p* orbital. To ionize the Xe atom, an electron must be removed from a 5*p* orbital. The 5*p* is a higher energy level and is farther from the nucleus than 2*p*, hence it takes less energy to remove an electron from Xe.**

(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:

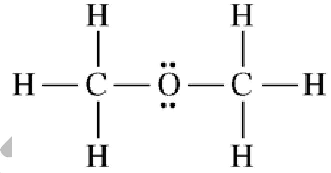
(i) The geometric shape of the XeO_3 molecule **Trigonal pyramidal**

(ii) The hybridization of the valence orbitals of xenon in XeF_4 **d^2sp^3**

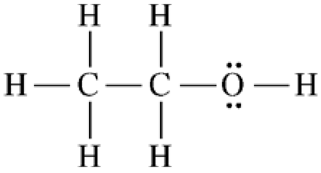
(f) Predict whether the XeO_3 molecule is polar or nonpolar. Justify your prediction. **The XeO_3 molecule would be polar because it contains three polar Xe –O bonds that are asymmetrically arranged around the central Xe atom (i.e., the bond dipoles do not cancel but add to a net molecular dipole with the Xe atom at the positive end).**

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

The intermolecular forces of attraction among molecules of dimethyl ether consist of London (dispersion) forces and weak dipole-dipole interactions. In addition to London forces and dipole-dipole interactions that are comparable in strength to those in dimethyl ether, ethanol can form hydrogen bonds between the H of one molecule and the O of a nearby ethanol molecule. Hydrogen bonds are particularly strong intermolecular forces, so they require more energy to overcome during the boiling process. As a result, a higher temperature is needed to boil ethanol than is needed to boil dimethyl ether

(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. **In the solid phase, SO_2 consists of discrete molecules with dipole-dipole and London (dispersion) forces among the molecules. These forces are relatively weak and are easily overcome at a relatively low temperature, consistent with the low melting point of SO_2. In solid SiO_2, a network of Si and O atoms, linked by strong covalent bonds, exists. These covalent bonds are much stronger than typical intermolecular interactions, so very high temperatures are needed to overcome the covalent bonds in SiO_2. This is consistent with the very high melting point for SiO_2.**

(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.

The intermolecular forces in liquid Cl_2 are London (dispersion) forces, whereas the intermolecular forces in liquid HCl consist of London forces and dipole-dipole interactions. Since the boiling point of Cl_2 is higher than the boiling point of HCl, the London forces among Cl_2 molecules must be greater than the London and dipole-dipole forces among HCl molecules. The greater strength of the London forces between Cl_2 molecules occurs because Cl_2 has more electrons than HCl, and the strength of the London interaction is proportional to the total number of electrons.

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. **4.02×10^{-19} J**

(c) Calculate the longest wavelength of light, in meters, that can supply the energy per photon necessary to break the Cl–Cl bond. **4.9×10^{-7} m**

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.

The total bond energy of the reactants is larger. Reaction Y is endothermic ($\Delta H_{298} = +41$ kJ mol−1 > 0), so there is a net input of energy as the reaction occurs. Thus, the total energy required to break the bonds in the reactants must be greater than the total energy released when the bonds are formed in the products.

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., $1s^2 2s^2 \ldots$) of each species.

S : $1s^2 2s^2 2p^6 3s^2 3p^4$

S^{2-} : $1s^2 2s^2 2p^6 3s^2 3p^6$

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

The nuclear charge is the same for both species, but the eight valence electrons in the sulfide ion experience a greater amount of electron-electron repulsion than do the six valence electrons in the neutral sulfur atom. This extra repulsion in the sulfide ion increases the average distance between the valence electrons, so the electron cloud around the sulfide ion has the greater radius.

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

The sulfur atom would be attracted into a magnetic field. Sulfur has two unpaired *p* electrons, which results in a net magnetic moment for the atom. This net magnetic moment would interact with an external magnetic field, causing a net attraction into the field. The sulfide ion would not be attracted into a magnetic field because all the electrons in the species are paired, meaning that their individual magnetic moments would cancel each other.

(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. **It requires less energy to remove an electron from a sulfide ion than from an argon atom. A valence electron in the sulfide ion is less attracted to the nucleus (charge +16) than is a valence electron in the argon atom (charge +18).**

(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?

The atomic orbitals involved in bonding with the H atoms in H_2S are *p* (specifically, 3*p*) orbitals. The three *p* orbitals are mutually perpendicular (i.e., at 90°) to one another.

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

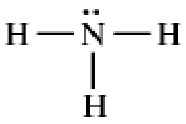
(i) Compare the strength of the London (dispersion) forces in liquid H_2S to the strength of the London (dispersion) forces in liquid H_2O. Explain.

The strength of the London forces in liquid H_2S is greater than that of the London forces in liquid H_2O. The electron cloud of H_2S has more electrons and is thus more polarizable than the electron cloud of the H_2O molecule.

(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.
The strength of the dipole-dipole forces in liquid H_2S is weaker than that of the dipole-dipole forces in liquid H_2O. The net dipole moment of the H_2S molecule is less than that of the H_2O molecule. This results from the lesser polarity of the H–S bond compared with that of the H–O bond (S is less electronegative than O).

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 $\Delta H°_{298}$ 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. **ΔH_{298} = Σ (bond energy of the reactants) − Σ (bond energy of the products). Based on the equation, for ΔH_{298} to be negative, the total bond energy of the products must be larger than the total bond energy of the reactants. OR More energy is released as product bonds are formed than is absorbed as reactant bonds are broken.**