1980

(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.

<table>
<thead>
<tr>
<th>Quantum #</th>
<th>4s electrons</th>
<th>4s electron</th>
<th>4p electron</th>
<th>4p electron</th>
<th>4p electron</th>
</tr>
</thead>
<tbody>
<tr>
<td>n</td>
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<td>4</td>
<td>4</td>
<td>4</td>
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</tr>
<tr>
<td>l</td>
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<td>0</td>
<td>1</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>m_l</td>
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<td>0</td>
<td>-1</td>
<td>0</td>
<td>1</td>
</tr>
<tr>
<td>m_s</td>
<td>+1/2</td>
<td>-1/2</td>
<td>+1/2</td>
<td>+1/2</td>
<td>+1/2</td>
</tr>
</tbody>
</table>

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

It would be paramagnetic because it has unpaired electrons in the 4p subshell.

(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\(_3\)As, AsCl\(_3\), and AsF\(_5\).

\[ \rightarrow \text{Na}_3\text{As}: \text{An ionic bond would form between sodium & arsenic, each sodium will transfer one electron to the Arsenic atom creating a fulfilled 4p orbital} \]

\[ \rightarrow \text{AsCl}_3: \text{A covalent bond would form between arsenic and chlorine atoms and would fulfill the remaining 3 open spaces in the 4p orbital.} \]

\[ \rightarrow \text{AsF}_5: \text{A covalent bond would form between arsenic and fluorine consequently because Arsenic is in period 4, it can have an expanded octet after the three electrons are filled in the 4p orbital} \]

\[ \rightarrow \text{We will review this concept more in the next Big Idea - Bonding} \]

1981

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

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

Quantized energy levels or discrete energies or wave properties of electron produce discrete energy states in a gas.

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

The excited state atoms can relax to several lower energy states (see diagram in c). Each final state energy level produces a separate series.
(c) Draw an electronic energy level diagram for the hydrogen atom and indicate on it the transition corresponding to the line of lowest frequency in the Balmer series.

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

- Emission spectra are photons emitted from excited state systems as they drop to lower energy states.
- Absorption spectra result from the absorption of electromagnetic radiation. Electrons are excited to a higher energy state.
- Hydrogen atoms are in the lowest electronic energy state at 25 °C (n = 1) so absorptions will be n = 1 to n = 2,3,4, etc.

1987
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.
   (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 = 0.99Å; Cl− ion = 1.81Å)

   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 for Mg: 1st = 7.6 ev; 2nd = 14 ev; 3rd = 80 ev)

The 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

Two important concepts that relate to the behavior of electrons in atom systems are the Heisenberg uncertainty principle and the wave-particle duality of matter.

(a) State the Heisenberg uncertainty principle as it related to the 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.
- The more exactly the position of a particle is known, the less exactly the position or velocity of the particle is known.
- OR you could describe Heisenberg’s uncertainty equation

\[ \Delta x \Delta p > \frac{\hbar}{4\pi} \]

\( \Delta x = \text{Uncertainty of Position} \)

\( \Delta p = \text{Uncertainty of Momentum} \)

(b) What aspect of the Bohr theory of the atom is considered unsatisfactory as a result of the Heisenberg uncertainty principle? NOTE: ALL 3 WERE ACCEPTABLE ANSWERS ACCORDING TO COLLEGEBOARD.ORG

- Bohr postulated that the electron in a H atom travels about the nucleus in a circular orbit and has a fixed angular momentum. With a fixed radius of orbit and a fixed momentum (or energy), and violates the uncertainty principle.
- The wavelength of a particle is given by the deBroglie relation: [gamma] = h/mv. For masses of macroscopic objects, h/m is so small for any v that [gamma] is so small as to be undetectable. For an electron, m is so small that h/mv yields a detectable [gamma].
- The product of the uncertainties in position and velocity depends on h/m and since h is so small (h = 6.626 \times 10^{-34} J \cdot s) unless m is very small, as for an electron, the product of the uncertainties is too small to be detected.
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 therefore the nuclear charge is increasing with stronger attraction for electrons and increases the IE (making it more difficult to remove the electrons from the element).

(b) The first ionization energy of B is lower than that of Be. The electrons in Be are 2s electrons (and the 2s orbital is completely filled) whereas in Be the electron is a 2p orbital (only one electron) therefore it is easier to remove the one electron even though the 2p electron is higher in energy than the 2s electrons.

(c) The first ionization energy of O is lower than that of N. The electron configuration for O is 2s², 2p⁴ and N is 2s², 2p³. The electron in O is paired with another electron in the same orbital whereas in N each electron occupies an orbital (Remember a half-filled orbital is MORE favorable) therefore the ionization energy is less for O because of the repulsion between two electrons in the same orbital (It would rather lose the one electron from the oxygen atom than lose an electron from one of the half-filled orbitals in N).

(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 that of both Li and Ne because the electron removed comes from 3s and is farther away from the nucleus therefore the electron is held less tightly by the nucleus. (Note: will Li also has one electron in the 2s orbital but it is closer to the nucleus and therefore held more tightly than the 3s electron in Na. Ne has a full orbital shell therefore it is more difficult to remove an electron.)
Account for each of the following in terms of principles of atom 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.

Na: [Ne] 3s
Mg: [Ne] 3s

For the 1st ionization energy of Na, it loses the 3s\(^{1}\) electron resulting in a full octet into the noble gas configuration of Ne therefore the second ionization energy would be high because it is extremely difficult to remove an electron from a full subshell. Whereas, for Mg the first ionization loses one electron from 3s resulting in the electron configuration of [Ne] 3s\(^{1}\) and the second ionization energy would be less because it is easier to lose the second 3s electron versus a 2p electron from Na.

(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.

Can discuss the following to get credit for the question:

- shielding differences
- energy differences between the elements
- the difference in the number of protons/electrons

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

Nickel has unpaired electrons-paramagnetic
Zinc has paired electrons/diamagnetic

(d) Phosphorus forms the fluorides PF\(_{3}\) and PF\(_{5}\), whereas nitrogen forms only NF\(_{3}\).

Expanded octet can occur in elements in period 3 or higher because there are available d orbitals for bonding whereas Nitrogen is too small to accommodate 5 fluorines or bonding sites.

1994 — SKIP QUESTION B FOR NOW

Use principles of atomic structure and/or chemical bonding to answer 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 then Ca. The outermost electrons 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\(\text(s)\) is -3,460 kilojoules per mole; the lattice energy for K\(_{2}\)O\(\text(s)\) is -2,240 kilojoules per mole. Account for this difference. Ionization Energy (kJ/mol)