1980

(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.

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?
(b) Account for the existence of several series of lines in the spectrum. What quantity distinguishes one series of lines from another?
(c) Draw an electronic energy level diagram for the hydrogen atom and indicate on it the transition corresponding to the line of lowest frequency in the Balmer series.
(d) What is the difference between an emission spectrum and an absorption spectrum? Explain why the absorption spectrum of atomic hydrogen at room temperature has only the lines of the Lyman series.

1987

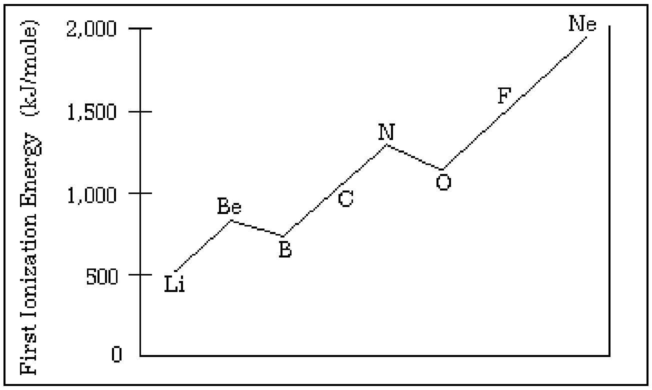
Use the details of modern atomic theory to explain each of the following experimental observations.
(a) Within a family such as the alkali metals, the ionic radius increases as the atomic number increases.
(b) The radius of the chlorine atom is smaller than the radius of the chloride ion, Cl^-. (Radii : Cl atom = 0.99Å; Cl^- ion = 1.81Å)
(c) The first ionization energy of aluminium is lower than the first ionization energy of magnesium. (First ionization energies: $_{12}Mg$ = 7.6 ev; $_{13}Al$ = 6.0 ev)
(d) For magnesium, the difference between the second and third ionization energies is much larger than the difference between the first and second ionization energies. (Ionization energies for Mg: 1^{st} = 7.6 ev; 2^{nd} = 14 ev; 3^{rd} = 80 ev)

1987

Two important concepts that relate to the behaviour 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.
(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 behaviour of macroscopic objects, but it is very significant when describing the behaviour of electrons.

1990

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.

1993

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.
(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 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 D

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.

(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/mol)	
	First	Second
K	419	3,050
Ca	590	1,140

(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.

1997

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

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.

2000

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.