The Advanced Placement Examination in Chemistry

Part I – Multiple Choice Questions
Part II – Free Response Questions
Selected Questions from 1970 to 2010

Atomic Theory and Periodicity

Part I

1984

1. Which of the following elements’ atoms forms monatomic ions with 2– charge in solutions?
   (A) F
   (B) S
   (C) Mg
   (D) Ar
   (E) Mn

19. Which of the following represents a pair of isotopes?

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Mass Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>(A) I. 6</td>
<td>14</td>
</tr>
<tr>
<td></td>
<td>II. 7</td>
</tr>
<tr>
<td>(B) I. 6</td>
<td>7</td>
</tr>
<tr>
<td></td>
<td>II. 14</td>
</tr>
<tr>
<td>(C) I. 6</td>
<td>14</td>
</tr>
<tr>
<td></td>
<td>II. 14</td>
</tr>
<tr>
<td>(D) I. 7</td>
<td>13</td>
</tr>
<tr>
<td></td>
<td>II. 14</td>
</tr>
<tr>
<td>(E) I. 8</td>
<td>10</td>
</tr>
<tr>
<td></td>
<td>II. 16</td>
</tr>
</tbody>
</table>

22. Atoms of an element, X, have the electronic configuration shown above. The compound most likely formed with magnesium, Mg, is
   (A) MgX
   (B) Mg₂X
   (C) MgX₂
   (D) MgX₃
   (E) Mg₃X₂

43. The elements in which of the following have most nearly the same atomic radius?
   (A) Be, B, C, N
   (B) Ne, Ar, Kr, Xe
   (C) Mg, Ca, Sr, Ba
   (D) C, P, Se, I
   (E) Cr, Mn, Fe, Co

58. Which of the following represents the ground state electron configuration for the Mn³⁺ ion? (Atomic number Mn = 25)
   (A) 1s²2s²2p⁶3s²3p⁶3d⁴
   (B) 1s²2s²2p⁶3s²3p⁶3d⁶4s²
70. One of the outermost electrons in a strontium atom in the ground state can be described by which of the following sets of four quantum numbers?
   (A) 5, 2, 0, ½
   (B) 5, 1, 1, ½
   (C) 5, 1, 0, ½
   (D) 5, 0, 1, ½
   (E) 5, 0, 0, ½

1989

Questions 1–3

(A) O
(B) La
(C) Rb
(D) Mg
(E) N

1. What is the most electronegative element?
2. Which element exhibits the greatest number of different oxidation states?
3. Which of the elements above has the smallest ionic radius for its most commonly found ion?

Questions 4–7

(A) $1s^22s^22p^63s^23p^5$
(B) $1s^22s^22p^63s^23p^6$
(C) $1s^22s^22p^62d^{10}3s^23p^6$
(D) $1s^22s^22p^63s^23p^63d^5$
(E) $1s^22s^22p^63s^23p^63d^44s^2$

4. An impossible electronic configuration
5. The ground–state configuration for the atoms of a transition element
6. The ground–state configuration of a negative ion of a halogen
7. The ground–state configuration of a common ion of an alkaline earth element

33. Which of the following conclusions can be drawn from J. J. Thomson’s cathode ray experiments?
   (A) Atoms contain electrons.
   (B) Practically all the mass of an atom is contained in its nucleus.
   (C) Atoms contain protons, neutrons, and electrons.
   (D) Atoms have a positively charged nucleus surrounded by an electron cloud.
   (E) No two electrons in one atom can have the same four quantum numbers.

64. A solution is known to contain an inorganic salt of one of the following elements. The solution is colorless. The solution contains a salt of
   (A) Cu
   (B) Mn
   (C) Fe
   (D) Ni
   (E) Zn
1994

Questions 1–4

(A) Heisenberg uncertainty principle
(B) Pauli exclusion principle
(C) Hund’s rule (principle of maximum multiplicity)
(D) Shielding effect
(E) Wave nature of matter

1. Can be used to predict that a gaseous carbon atom in its ground state is paramagnetic
2. Explains the experimental phenomenon of electron diffraction
3. Indicates that an atomic orbital can hold no more than two electrons
4. Predicts that it is impossible to determine simultaneously the exact position and the exact velocity of an electron

27. Which of the following sets of quantum numbers \((n, l, m_l, m_s)\) best describes the valence electron of highest energy in a ground–state gallium atom (atomic number 31)?
   (A) 4, 0, 0, \(\frac{1}{2}\)
   (B) 4, 0, 1, \(\frac{1}{2}\)
   (C) 4, 1, 1, \(\frac{1}{2}\)
   (D) 4, 1, 2, \(\frac{1}{2}\)
   (E) 4, 2, 0, \(\frac{1}{2}\)

46. Which of the following solids dissolves in water to form a colorless solution?
   (A) CrCl\(_3\)
   (B) FeCl\(_3\)
   (C) CoCl\(_2\)
   (D) CuCl\(_2\)
   (E) ZnCl\(_2\)

1999

Questions 5–8 refer to atoms for which the occupied atomic orbitals are shown below.

(A) 1\(s\) 2\(s\) 
(B) 1\(s\) 2\(s\) 
(C) 1\(s\) 2\(s\) 2\(p\) 
(D) 1\(s\) 2\(s\) 2\(p\) 
(E) [Ar] 3\(s\) 3\(d\)

5. Represents an atom that is chemically unreactive
6. Represents an atom in an excited state
7. Represents an atom that has four valence electrons
8. Represents an atom of a transition metal

21. When a solution of sodium chloride is vaporized in a flame, the color of the flame is
   (A) blue
   (B) yellow
   (C) green
   (D) violet
   (E) white
### Atomic Theory and Periodicity

#### Ionization Energies for element $X$ (kJ mol$^{-1}$)

<table>
<thead>
<tr>
<th></th>
<th>First</th>
<th>Second</th>
<th>Third</th>
<th>Fourth</th>
<th>Fifth</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>580</td>
<td>1,815</td>
<td>2,740</td>
<td>11,600</td>
<td>14,800</td>
</tr>
</tbody>
</table>

37. The ionization energies for element $X$ are listed in the table above. On the basis of the data, element $X$ is most likely to be
   (A) Na  
   (B) Mg  
   (C) Al  
   (D) Si  
   (E) P

50. In the periodic table, as the atomic number increases from 11 to 17, what happens to the atomic radius?
   (A) It remains constant.  
   (B) It increases only.  
   (C) It increases, then decreases.  
   (D) It decreases only.  
   (E) It decreases, then increases.

51. Which of the following is a correct interpretation of the results of Rutherford’s experiments in which gold atoms were bombarded with alpha particles?
   (A) Atoms have equal numbers of positive and negative charges.  
   (B) Electrons in atoms are arranged in shells.  
   (C) Neutrons are at the center of an atom.  
   (D) Neutrons and protons in atoms have nearly equal mass.  
   (E) The positive charge of an atom is concentrated in a small region.

2002

Questions 1-2

Consider atoms of the following elements. Assume that the atoms are in the ground state
   (A) S  
   (B) Ca  
   (C) Ga  
   (D) Sb  
   (E) Br

1. The atom that contains exactly two unpaired electrons.
2. The atom that contains only one electron in the highest occupied energy sublevel.

17. In which of the following groups are the three species isoelectronic; i.e., have the same number of electrons?
   (A) $S^{2-}$, $K^+$, $Ca^{2+}$  
   (B) $Sc$, $Ti$, $V^{2+}$  
   (C) $O^{2-}$, $S^{2-}$, $Cl^-$  
   (D) $Mg^{2+}$, $Ca^{2+}$, $Sr^{2+}$  
   (E) $Cs$, $Ba^{2+}$, $La^{3+}$

44. Which of the following properties generally decreases across the periodic table from sodium to chlorine?
   (A) First ionization energy
(B) Atomic mass  
(C) Electronegativity  
(D) Maximum value of oxidation number  
(E) Atomic radius  

46. The effective nuclear charge experienced by the outermost electron of Na is different than the effective nuclear charge experienced by the outermost electron of Ne. This difference best accounts for which of the following?  
(A) Na has a greater density at standard conditions than Ne.  
(B) Na has a lower first ionization energy than Ne.  
(C) Na has a higher melting point than Ne.  
(D) Na has a higher neutron-to-proton ratio than Ne.  
(E) Na has fewer naturally occurring isotopes than Ne.  

59. All of the halogens in their elemental form at 25°C and 1 atm are  
(A) conductors of electricity  
(B) diatomic molecules  
(C) odorless  
(D) colorless  
(E) gases  

2008  
3. Which of the following types of elements are most likely to form anions?  
(A) Noble gases  
(B) Alkali metals  
(C) Halogens  
(D) Transition elements  
(E) Actinides  

19. Which of the following ions has the same number of electrons as Br\(^{-}\)?  
(A) Ca\(^{2+}\)  
(B) K\(^{+}\)  
(C) Sr\(^{2+}\)  
(D) I\(^{-}\)  
(E) Cl\(^{-}\)  

21. Of the following electron configurations of neutral atoms, which represents an atom in an excited state?  
(A) 1s\(^2\)2s\(^2\)2p\(^5\)  
(B) 1s\(^2\)2s\(^2\)2p\(^3\)3s\(^2\)  
(C) 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^1\)  
(D) 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^2\)  
(E) 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^5\)  

23. The oxidation state that is common to aqueous ions of Fe, Mn, and Zn is  
(A) +1  
(B) +2  
(C) +3  
(D) +4  
(E) +5
24. Which of the following shows the correct number of protons, neutrons, and electrons in a neutral cesium-134 atom?

<table>
<thead>
<tr>
<th></th>
<th>Protons</th>
<th>Neutrons</th>
<th>Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>(A)</td>
<td>55</td>
<td>55</td>
<td>55</td>
</tr>
<tr>
<td>(B)</td>
<td>55</td>
<td>79</td>
<td>55</td>
</tr>
<tr>
<td>(C)</td>
<td>55</td>
<td>79</td>
<td>79</td>
</tr>
<tr>
<td>(D)</td>
<td>79</td>
<td>55</td>
<td>79</td>
</tr>
<tr>
<td>(E)</td>
<td>134</td>
<td>55</td>
<td>134</td>
</tr>
</tbody>
</table>

63. Which of the following best helps to account for the fact that the F\(^{-}\) ion is smaller than the O\(^{2-}\) ion?

(A) F\(^{-}\) has a larger nuclear mass than O\(^{2-}\) has.
(B) F\(^{-}\) has a larger nuclear charge than O\(^{2-}\) has.
(C) F\(^{-}\) has more electrons than O\(^{2-}\) has.
(D) F\(^{-}\) is more electronegative than O\(^{2-}\) is.
(E) F\(^{-}\) is more polarizable than O\(^{2-}\) is.

Part II

1970
What is meant by the lanthanide contraction? Account for this phenomenon. Give two examples of its consequences.

1971
There is a greater variation between the properties (both chemical and physical) of the first and second of a group or family in the periodic table than between the properties of the second and third members of the group. Consider as examples either the group containing nitrogen or the one containing oxygen. Select three properties and discuss the variation of these properties to illustrate the generalization expressed in the first sentence of the question.

1972
Consider the following melting points in degrees Celsius:

<table>
<thead>
<tr>
<th>Alkali metals</th>
<th>Halogens</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Li 181°</td>
<td>F(_2) -119°</td>
<td></td>
</tr>
<tr>
<td>Na 98°</td>
<td>Cl(_2) -101°</td>
<td></td>
</tr>
<tr>
<td>K  63°</td>
<td>Br(_2) -7°</td>
<td></td>
</tr>
<tr>
<td>Rb 39°</td>
<td>I(_2) +104°</td>
<td></td>
</tr>
<tr>
<td>Cs 29°</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

(a) Account for the trend in the melting points of the alkali metals.
(b) Account for the trend in the melting points of the halogens.

1973

| Li 124 | 1.34 |
| Be 215 | 0.90 |
| B 191  | 0.82 |
| C 260  | 0.77 |
| N 336  | 0.75 |
| O 314  | 0.73 |
| F 402  | 0.72 |
The covalent radii decrease regularly from Li to F, whereas the first ionization energies do not. For the ionization energies, show how currently accepted theoretical concepts can be used to explain the general trend and the two discontinuities.

1976
\[ \text{M(s)} + \text{Cl}_2(g) \rightarrow \text{MCl}_2(s) \]
The reaction of a metal with chlorine proceeds as indicated above. Indicate, with reasons for your answers, the effect of the following factors on the heat of reaction for this reaction.
(a) A large radius versus a small radius for M²⁺
(b) A high ionization energy versus a low ionization energy for M.

1977
The electron affinities of five elements are given below.
\begin{align*}
13\text{Al} & \quad 12 \text{ kcal/mole} \\
14\text{Si} & \quad 32 \text{ kcal/mole} \\
15\text{P} & \quad 17 \text{ kcal/mole} \\
16\text{S} & \quad 48 \text{ kcal/mole} \\
17\text{Cl} & \quad 87 \text{ kcal/mole}
\end{align*}
Define the term “electron affinity” of an atom. For the elements listed above, explain the observed trend with the increase in atomic number. Account for the discontinuity that occurs at phosphorus.

1978
The postulates of the Bohr model of the hydrogen atom can be stated as follows:
(I) The electron can exist only in discrete states each with a definite energy.
(II) The electron can exist only in certain circular orbits.
(III) The angular momentum of the electron is \( nh/2\pi \) where \( n \) is any positive integer.
(IV) Radiation is emitted by the atom only when an electron makes a transition from a state of higher energy to one of lower energy.
(a) State whether each of these postulates is currently considered to be correct, according to the wave mechanical description of the hydrogen atom.
(b) Give the wave mechanical description that has replaced one of the postulates now considered to be incorrect.

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) Explain how the electron configuration of the arsenic atom in the ground state is consistent with the existence of the following known compounds: \( \text{Na}_3\text{As} \), \( \text{AsCl}_3 \), and \( \text{AsF}_5 \).

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?
1982

The values of the first three ionization energies \((I_1, I_2, I_3)\) for magnesium and argon are as follows:

<table>
<thead>
<tr>
<th></th>
<th>I_1 (kJ/mol)</th>
<th>I_2 (kJ/mol)</th>
<th>I_3 (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg</td>
<td>735</td>
<td>1443</td>
<td>7730</td>
</tr>
<tr>
<td>Ar</td>
<td>1525</td>
<td>2665</td>
<td>3945</td>
</tr>
</tbody>
</table>

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

<table>
<thead>
<tr>
<th></th>
<th>I_1 (kJ/mol)</th>
<th>I_2 (kJ/mol)</th>
<th>I_3 (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Q</td>
<td>496</td>
<td>4568</td>
<td>6920</td>
</tr>
</tbody>
</table>

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.

1984

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.

1987 (1)

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, \(\text{Cl}^-\). (Radii: \(\text{Cl} \) atom = 0.99 Å; \(\text{Cl}^-\) ion = 1.81 Å)
(c) The first ionization energy of aluminum is lower than the first ionization energy of magnesium. (First ionization energies: \(^{13}\text{Mg} = 7.6 \text{ eV}; ^{13}\text{Al} = 6.0 \text{ 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\text{st} = 7.6 \text{ eV}; ^2\text{nd} = 14 \text{ eV}; ^3\text{rd} = 80 \text{ eV})

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.

1994
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 nanometers; the radius of the Ca\(^{2+}\) ion is 0.099 nanometers. Account for this difference.

<table>
<thead>
<tr>
<th>Ionization Energy (kJ/mol)</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>First</td>
<td>Second</td>
</tr>
<tr>
<td>K</td>
<td>419</td>
</tr>
<tr>
<td>Ca</td>
<td>590</td>
</tr>
</tbody>
</table>

(b) Explain the difference between Ca and K in regard to
(i) their first ionization energies,
(ii) their second ionization energies.

(c) 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 of 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., 1$s^2$2$s^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).

2005

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

<table>
<thead>
<tr>
<th>Element</th>
<th>First Ionization Energy (kJ mol$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Si</td>
<td>786</td>
</tr>
<tr>
<td>P</td>
<td>1012</td>
</tr>
<tr>
<td>Cl</td>
<td>1251</td>
</tr>
</tbody>
</table>

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

(b) A certain element has two stable isotopes. The mass of one of the isotopes is 62.93 amu and the mass of the other isotope is 64.93 amu.
   (i) Identify the element. Justify your answer.
   (ii) Which isotope is more abundant? Justify your answer.

2008

Using principles of atomic and molecular structure and the information in the table below, answer the following questions about atomic fluorine, oxygen, and xenon.
<table>
<thead>
<tr>
<th>Atom</th>
<th>First Ionization Energy (kJ mol(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>F</td>
<td>1,681.0</td>
</tr>
<tr>
<td>O</td>
<td>1,313.9</td>
</tr>
<tr>
<td>Xe</td>
<td>?</td>
</tr>
</tbody>
</table>

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

(b) Account for the fact that the first ionization energy of atomic fluorine is greater than that of atomic oxygen.

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

2009 (1)
Initiating most reactions involving chlorine gas involves breaking the Cl–Cl bond, which has a bond energy of 242 kJ mol\(^{-1}\).

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

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

2009 (2)
Consider the two chemical species S and S\(^{2-}\).

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

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

(c) 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.

2010
Respond to the following statements and questions that relate to zinc and its ion.

(a) Write the complete electron configuration (e.g., 1s\(^2\)2s\(^2\) . . .) for Zn\(^{2+}\).

(b) Which species, Zn or Zn\(^{2+}\), has the greater ionization energy? Justify your answer.