Atomic Origins: Chapter Problems

Atomic Masses

Class work
1. The atomic model has evolved significantly since Dalton’s original atomic theory. Discuss evidence that supports or disproves Dalton’s postulates.
2. Describe the discovery of the electron’s charge and mass.
3. Calculate the Coulombic force between a proton and an electron in a hydrogen atom (r = 53pm).
4. Describe the plum pudding model of the atom and experiments that supported or disproved this model.
5. Distinguish between atomic number, atomic mass number, and atomic mass.
6. Silicon is a metalloid used in the computer industry. It consists of three isotopes: Si – 28 (27.9769); Si – 29 (28.9765); Si – 30 (29.9794). Which isotope is most abundant? Explain your reasoning.
7. If 3.09% of all silicon atoms in a sample are Si-30, what are the abundances of the other two isotopes?
8. Silver consists of two stable isotopes, one with a mass of 106.90509 and an abundance of 51.84%. What is the abundance and mass of the other isotope?
9. Explain how a mass spectrometer functions.
10. The mass reported on the periodic table for chlorine is 35.45 u. Why, when a sample of chlorine gas is examined with a mass spectrometer, there is no peak in the spectrum with a mass of 35.45 u?
11. A mass spectrum can report the data from a mass spectrometer on an intensity vs. mass graph or a percent abundance vs. mass graph. Explain the difference between these two graphs.
12. Below is the mass spectrum of zirconium (Zr). Determine the average atomic mass of Zr. Justify with calculations.
Below can be found the mass spectrum of a pure element. Use this spectrum to answer questions 13-14.

13. Using a ruler, determine the % abundance of each isotope.
14. Calculate the average atomic mass and identify the element.
15. Draw the mass spectrum (percent abundance vs. mass) for silicon. Clearly label axes, including units.
16. Draw the mass spectrum (intensity vs. mass) for oxygen. Clearly label axes, including units.

**Homework**

17. The charge/mass ratio of an electron was determined by Thompson's cathode ray experiments. How would this charge/mass ratio differ if Thompson used anode rays?
18. Calculate the Coulombic force between the 2 electrons and 2 protons in a helium atom ($r_{He} = 31\text{pm}$).
19. Compare the Coulombic forces between the electrons and protons with an oxygen atom to those forces within a nitrogen atom. Predict how this will affect the structure and properties of these atoms.
20. In 1935, James Chadwick was awarded the Noble Prize for discovering the neutron. How did this particle change our understanding of the atom?
21. How is the Rutherford nuclear model different than the plum pudding model and what experimental evidence supported his model?
22. What do isotopes of an element have in common? How do they differ?
23. Chromium has four stable isotopes. Three of these are: Cr-50 with a percent abundance of 4.35%; Cr-52 with a percent abundance of 83.8%; Cr-53 with a percent abundance of 9.5%. What is the percent abundance and mass number of the fourth isotope?
24. Oxygen has three stable isotopes: O-16 (99.75%); O-17 (0.04%); O-18 (0.21%). What are the masses of these three isotopes (in u)?
25. Why must the samples placed in a mass spectrometer be ionized?
Below is the mass spectrum of a pure element. Use this spectrum to answer questions 26-27.

26. What should the y-axis of this graph be labeled? Justify your reasoning
27. Determine the average atomic mass of this element and identify the element.
28. Draw the mass spectrum (percent abundance vs. mass) for copper. Cu-64 has a percent abundance of 69.15% and Cu-65 has a percent abundance 30.85%. Clearly label axes, including units.
29. Draw the mass spectrum (intensity vs. mass) for chromium.
30. Radiocarbon dating is a technique that allows scientists to determine the age of a substance based on the ratio of carbon-14 to carbon-12 in the substance. Carbon-14 is known to decay to Nitrogen-14 with a half-life of 5730 years. How can mass spectrometry be useful to the radiocarbon dating process?

Atomic Modeling
Class work
31. What evidence contradicted the Rutherford model of the atom?
32. How did Niels Bohr explain the existence of spectral lines?
33. When lithium is vaporized and energized, a spectral line with wavelength of roughly 698 nm is produced. Is this more or less energetic than light with a wavelength of 650 nm? Explain.
34. How much energy is emitted by the photon of light with a wavelength of 698 nm?
35. Which electronic transition within a lithium atom would emit more energy, 2→1 or 3→2? Justify your answer.
36. What color spectral line would be created if the energy difference between orbits in the atom is $4.85 \times 10^{-19}$ J?
37. How many orbits would Bohr predict he would find in a ground state atom of magnesium?
38. Describe how photoelectron spectroscopy works.
Use the modified PES spectra for helium, neon, and krypton below to answer questions 39-42.

39. Explain how helium (despite having two electrons) demonstrates only 1 peak and how this supports the Bohr model of the atom.

40. The PES spectra for neon indicates there are how many discrete orbits where electrons could exist? For Krypton?

41. Were the number of orbits for Ne and Kr greater or less than Bohr had predicted?

42. What is the binding energy of the peak in the neon spectra that is produced by the removal of electrons in the orbit closest to the nucleus? Justify your reasoning.

43. Sketch the PES spectrum of lithium indicating the relative intensities and binding energies of the peaks.

44. How did de Broglie explain the stability of the Bohr orbits and explain why this prompted the development of the quantum model of the atom?

45. Both the Bohr and Quantum model predicted 8 electrons could fit in the second main energy level. Describe how the quantum model differs from the Bohr model in how these electrons are arranged in this second main energy level.

46. Explain how the quantum model was able to explain the PES spectra for boron shown below, ie: which orbitals represent which peaks.

47. Predict the number of peaks present in the PES spectrum of silicon. Compare their intensities to those of boron.

48. Sketch the PES spectrum of chlorine indicating the relative intensities and binding energies of the peaks.

49. Would the peak representing the 3d subshell lie to the left or the right of the peak for the 4s subshell in a PES spectrum of Cobalt?

50. How many unpaired electrons are in Cr?

51. Write the correct electron configuration for fluorine.
52. Write the correct electron configuration for Ti and Ti$^{2+}$
53. Write the correct electron configuration for N and N$^{3-}$.
54. How does PES data support the Aufbau principle?
55. How does an element in its excited state differ from an ion?
56. Write the correct electron configuration for Li$^*$.  
57. Write the correct electron configuration for Cs$^*$. 

**Homework**

58. What law of physics did Bohr violate in his explanation of atomic structure?
59. Distinguish between “ground states” and “excited states”.
60. Define “quantized”.
61. How are emission spectra and absorption spectra related?
62. The difference in energy between Bohr orbits $N=1$ and $N=3$ in a hydrogen atom is $1.93\times10^{-18}$ J. What wavelength (in nm) of light would need to be absorbed to promote an electron transition from $N=1$ to $N=3$?
63. How much energy is emitted by a photon of blue light (430 nm)? Would yellow light produce more or less energy?
64. Red light (700 nm) is produced when the energy of the photon hitting the atom is $2.84\times10^{-19}$ J. Determine the speed of the photon. Justify with calculations.
65. The photoelectric effect is the observation that many metals emit electrons when light shines on them. Discuss the implication of this effect in determining atomic structure. Does this effect have engineering implications?
66. What determines the intensity of the electron signal and the binding energy of a particular peak on a PES spectrum?

**Use the simplified PES spectrum of carbon below to answer questions 67-68.**

67. Discuss how this spectrum supports and/or contradicts the Bohr model of the atom.
68. How would the spectrum of oxygen compare to the spectrum of carbon?
69. Justify the following statement: The peak with the highest energy represents the electrons in the lowest atomic orbit.
70. What do the four quantum numbers represent?
71. Predict the differences between the PES spectrum of Ga and that of Fe. How many more peaks would the Ga PES spectra have compared to Fe? How much greater should the intensity of the 3d peak in Fe be compared to the 4s peak?
72. How does PES data support the Pauli exclusion principle?

Use the following website to answer questions 73-77
http://www.chem.arizona.edu/chemt/Flash/photoelectron.html
73. Why does boron’s lowest energy peak have less intensity than the higher energy peak at 1.36?
74. Why is calcium’s highest energy PES peak the same intensity as its lowest energy PES peak?
75. For the aluminum spectrum, which orbitals and how many electrons are present in the peaks observed with binding energies of: 151, 7.19, and 0.58?
76. Examine the PES spectra for sulfur and determine what evidence exists to support the quantum model assertion that all 3 orientations of the p orbital are degenerate – ie they are of the same energy.
77. For the element scandium, identify which orbital is responsible for the peaks observed with binding energies of: 48.5, 0.63, 0.77, and 433
78. Transition metals, like Sc, lose their 4s electrons before their 3d electrons. What evidence do you see in the PES spectra to support this observation?
79. Again, using evidence from the PES spectra, describe why Sc is unlikely to lose more than 3 electrons.
80. Write the correct electron configuration for the Sc^{2+} ion
81. Criticize the following statement: Hund’s rule dictates that electrons will maintain the maximum number of unpaired electrons.
82. How many unpaired electrons are in Sn^{4+} and Sn?
83. Write the correct electron configuration for iron.
84. Write the correct electron configuration for I and I^{-}.
85. Write the correct electron configuration for copper in an excited state.

The Periodic Table
Class work
86. Describe the arrangement of the periodic table.
87. An ion has 12 electrons. Can you determine its identity or placement on the periodic table? Explain.
88. An atom forms an ion which combines with the aluminum ion in a 3:1 ratio. Which group must this element belong to on the periodic table?
89. Write the correct shorthand electron configuration for magnesium.
90. Write the correct shorthand electron configuration for O^{2-}.
91. An element has an electron configuration of [X] ns^{1}. In which group does it belong?
92. An element has an electron configuration of [X] ns^{1}nd^{5}. What principle does this violate? Why does this occur?
93. Why are alkaline earth metals more stable than alkali metals?
94. The halogens are the most reactive group of non-metals. Propose a reason for this.
95. Give the shorthand electron configuration of copper.

Homework
96. What is the periodic law?
97. What is significant about Group 18 of the periodic table?
98. Write the correct shorthand electron configuration for selenium.
99. Write the correct shorthand electron configuration for Br*.
100. An element has an electron configuration of [X] ns²nd³. Where does this element belong on the periodic table?
101. Why is manganese more stable than iron?
102. Predict the most stable elements in the f-block.
103. Give the shorthand electron configuration of chromium.
104. Why isn’t chlorine’s electron configuration an exception?

**Periodic Trends**

**Class work**

105. Justify the following statement: As atomic number increases down a group, the atomic radius of the atoms will increase.
106. Boron, aluminum, and gallium are in group 13. Their atomic radii are (in pm) 85, 143, and 135. Why is the atomic radius of gallium an unexpected value? What can you discern about the shielding ability of the 3d orbital based on this data?

**Use the PES spectra for nitrogen and fluorine below to answer questions 107-109**

![PES spectra for nitrogen and fluorine](image)

107. Which of these two would have the smaller atomic radii? Justify your answer using data from the PES spectra.
108. What is responsible for the difference in atomic radii? (Shielding or nuclear charge) Justify your answer.
109. Why is fluorine’s 1s peak of a higher binding energy than the nitrogen’s 1s peak?
110. Magnesium tends to form ions with a 2+ charge. Would these ions have an ionic radii greater or less than the atomic radii of neutral magnesium? Justify your answer.
111. Rank the following atoms/ions in order of increasing atomic radii: Ga, Ge, Si
112. Rank the following atoms/ions in order of increasing atomic radii: P, S, Ar, P³⁻
113. Define electronegativity and provide an explanation as to why the noble gases do not have measurable electronegativity values.
114. Draw a graph that shows the general trend in electronegativity when plotted vs. the atomic number for the first 18 elements. Do not use specific electronegativity values but do ensure your points are internally consistent.
115. Rank the following elements in order of decreasing electronegativity: Ca, Al, Mg, P
116. For the following compounds, indicate which element would have the strongest pull on electrons and explain why: SF₆, H₂O, N₂O, CH₃Cl

117. What are three factors that influence ionization energy? Explain.

118. Explain the following observations: The first ionization energy of Mn is 717 kJ/mol while the first ionization energy of Cr is 656 kJ/mol.

Use the PES spectrum for the unknown element below to answer questions 119-123

![PES Spectrum]

119. Identify the type of orbital responsible for the peak at: 0.63, 0.77, and 3.24

120. Identify the element.

121. From which orbital would electrons be lost most easily? Justify your answer.

122. What would be the most common ionic charges for this element? Justify your answer.

123. Why would a 4+ ion be highly unlikely for this element? Justify your answer.

124. Do metallic elements typically have high or low ionization energies? Explain your answer.

125. Both bromine and copper are in the same period. Explain why bromine is a non-conductor of electricity while copper is an excellent conductor of electricity.

126. Draw a graph that shows the trend in metallic character for group 16 when plotted vs. atomic number. Do not use quantitative values for metallic character but make sure your data is internally consistent.

127. Oxygen and sulfur both form ions with a -2 charge, despite sulfur having 16 electrons to oxygen’s 8. Based on the periodic law, we might expect the next element to form a -2 ion to have 24 electrons, but instead we must wait for an element with 34 electrons to form a -2 anion. Why?

128. For each of the following cations, explain from which orbital the electrons were lost and how many. In addition, indicate which orbital electrons would be lost from first by listing that orbital first: Mn⁷⁺, Cr³⁺, K⁺, Ga³⁺, Sn⁴⁺, Cd²⁺, and Ag⁺

Homework

129. Francium has the largest atomic radii of any element (270 pm) on the periodic table. Explain how this is possible considering it has a very high nuclear charge.
130. If one were to graph the atomic radii versus atomic number, the graph would look like the one below. Explain the substantial increase in atomic radii from He to Li and the decrease in atomic radii from Na to Ar.

131. Explain why fluorine has a smaller atomic radii (42 pm) than oxygen (48 pm).

132. Which electrons (core or valence) are responsible for shielding the nuclear charge and how does this shielding influence the atomic radii?

133. We might normally define the atomic radius as being the distance between the nucleus and the outermost electron(s). Based on our understanding of the quantum model of the atom, why can this not be measured with certainty?

134. Magnesium tends to form ions with a 2+ charge. Would these ions have an ionic radii greater or less than the ionic radii of fluoride (F-)? Justify your answer.

135. Rank the following atoms/ions in order of decreasing atomic radii: Ca, Br, Br-, Ca2+

136. Rank the following atoms/ions in order of decreasing atomic radii: Cl, Cl-, Ar

137. Explain the general trend in electronegativity as the atomic number increases down a group AND provide a rationale for this trend.

138. Chlorine has 8 more protons in its nucleus than does fluorine. Provide a reason for its observed electronegativity (3.2) being lower than that of fluorine (4.0).

139. Rank the following elements in order of decreasing electronegativity: C, Si, S, N

140. Relate ionization energy to atomic radii and electronegativity.

141. Explain why the noble gases have the highest ionization energies in their periods.

142. Rank the following elements in order of increasing first ionization energy: Na, K, Sc.

143. Explain the difference between Ca and K in regard to their first ionization energies and their second ionization energies.

Use the binding energies from the PES peak for chlorine and aluminum's 3p electrons below to answer questions 137-138

<table>
<thead>
<tr>
<th>Element</th>
<th>Binding Energy of 3p electrons (x 10^{-19} J)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al</td>
<td>0.58</td>
</tr>
<tr>
<td>Cl</td>
<td>1.25</td>
</tr>
</tbody>
</table>

144. Explain why aluminum tends to form cations and chlorine tends to form anions?

145. Scandium (Sc) and Aluminum both form +3 cations. Explain how this is possible despite Al and Sc being in different groups:

146. Which of the processes below represents the following: the first ionization of magnesium? The second ionization of magnesium? The process requiring the most input of energy?
Mg → Mg⁺ + e⁻  
Mg⁺ + e⁻ → Mg  
Mg⁺ → Mg²⁺ + e⁻  
Mg²⁺ + e⁻ → Mg⁺

147. Lead and carbon reside in the same period yet lead demonstrates metallic properties while carbon does not. Explain.

148. Rank the following elements from most metallic to least metallic: Ga, Sr, Na, K, Fr.

149. Silver is in the same group as copper. Is it more or less metallic than copper? Explain.

150. Atoms within the same group tend to form ions with similar charges whereas ionic charges tend to vary significantly throughout a period. Explain this observation.

151. Lead generally forms ions with a +2 and a +4 charge. Which electrons were lost to form each?

152. What would be the expected ionic charges for Ge and Tl?

**Free Response**

1. Tungsten (W) has five naturally abundance isotopes: $^{180}$W (0.12%), $^{182}$W (26.50%), $^{183}$W (14.31%), $^{184}$W (30.64%), and $^{186}$W (28.43%).
   a. How many protons and neutrons would be found in each isotope’s nucleus?
   b. Explain how the percent abundance of each isotope would have been obtained.
   c. Draw the mass spectrum (percent abundance vs. mass) for this element.
   d. Write the shorthand electron configuration for this element.

2. Use the mass spectrum below to answer the following questions.

   ![Mass Spectrum Image]

   a. Redraw this mass spectrum as a percent abundance vs. mass graph.
   b. Determine the average atomic mass of this element. Identify the element.
   c. State the number of protons and neutrons in each isotope.
   d. How did the discovery of isotopes affect our understanding of atomic structure?
3. Using the following clues, identify the elements:
   a. Halogen with an atomic radii smaller than barium but larger than chlorine.
   b. Has three unpaired electrons in its 4d orbital.
   c. Transition metal with the highest ionization energy.
   d. Alkali metal with the smallest atomic radii.
   e. Transition metal with the largest atomic radii with a partially filled 4d orbital.
   f. Period 2 element with 2 unpaired electrons.
   g. Alkaline earth metal with the lowest 3rd ionization energy.

4. Group 16 elements are sometimes referred to as the chalcogens.
   a. Which element would be most metallic within this group? Justify your answer.
   b. When oxygen is combined with fluorine, which element would have the:
      i. Larger ionization energy? Justify your answer.
      ii. Smaller atomic radii? Justify your answer

5. The following questions pertain to the graph below:

   a. Explain why, in general, there is an increase in the first ionization energy from Li to Ne.
   b. Explain the non-linear nature of the graph.
   c. Predict how the first ionization energy of Na compares to those of Li and of Ne. Justify your prediction.
   d. Sketch PES spectra for Be and B. State the cations most commonly formed by these elements. Justify your answer based on the PES spectra.
6. The PES spectrum for lithium and helium can be seen below:

![Intensity diagram for Li and He with binding energies 0.52, 6.26, and 2.37 x 10^{-19} J]

a. Which orbitals are responsible for producing the peaks listed at binding energies:
   i. 0.52?
   ii. 6.26?
   iii. 2.37?

b. Why is less energy required to remove lithium’s 2s electron compared to helium’s 1s electrons?

c. Which element would have the highest:
   i. First ionization energy? Justify your answer.
   ii. Second ionization energy? Justify your answer.

d. Do these spectra support or contradict the Bohr model of the atom? Justify your answer.

7. The PES spectrum for sulfur is shown below:

![Intensity diagram for Sulfur with binding energies 1.0, 2.05, 16.5, 22.7, and 239 x 10^{-19} kJ]

a. Which peak corresponds to sulfur’s valence “p” orbital electrons?

b. Write the complete electron configuration for sulfur.

c. Would the same peak in Cl be expected to have a higher or lower binding energy compared to S? Justify your answer.

d. Compare the radial size of the sulfide ion to its parent atom.

e. Does this spectrum support or contradict the Bohr model of the atom? Justify your answer.
**Answers**

1. Dalton’s postulate regarding the indivisibility of atoms was disproved by the discovery of the subatomic particles and the nuclear reactions. His postulate regarding atoms of the same element being identical was disproved by the gold foil experiment. His postulate regarding different elements being different is supported by evidence that each element has a different number of protons. His postulate regarding elements not being changed in a chemical reaction is supported by the conservation laws, except in cases of nuclear reactions.

2. The electron was determined to be negative and its charge and mass ratio found through the cathode ray experiments. In these experiments, the cathode ray was deflected using electric and magnetic fields of known strength. The charge of the electron was determined through the oil drop experiment in which the charge distributed on oil drops was measured and determined to always occur in quanta of $1.6 \times 10^{-19}$. The charge/mass ratio was then used to determine the mass of a single electron.

3. $8.2 \times 10^{-8}$ N

4. The plum pudding model of the atom consisted of a positively charge sphere with electrons embedded in it. It was developed because it was support by the finding that protons were much more massive than electrons and by Coulomb’s law of attraction between oppositely charged particles. It was disproved by the gold foil experiment which demonstrated that the atom was mostly empty space with a small dense core.

5. The atomic number is the number of protons in the atom. The atomic mass number or just mass number is the number of nucleons in the atom. The atomic mass is the average mass of all isotopes of the element based on their percent abundance.

6. Si-28 is the most abundant isotope because the average atomic mass is 28.086 u.

7. Si-28 92.14%; Si-29 4.77%

8. 48.16%; 108.91 u

9. A mass spectrometer ionizes the sample and then subject it to a magnetic field. The greater the mass of the ion, the less it will be deflected, allowing the machine to record the mass/charge ratio of the sample.

10. 35.45 u is the average mass of the stable isotopes of this element. No individual chlorine atom has this mass and thus it is not recorded on the mass spec.

11. An intensity vs. mass records the most abundant isotope with an intensity of 100 and relates all other isotopes abundances to this value. The percent abundance vs. mass graph is produced when the intensity of each isotope is divided by the total of all the isotopic intensities.

12. 91.224 u

13. 196 = 0.15%, 198 = 9.97%, 199 = 16.9%, 200 = 23.10%, 201 = 13.2%, 202 = 29.9%, 204 = 6.9%

14. 200.59, Hg

15. Silicon
16. Oxygen

17. The ratio would be smaller because the mass of a proton is much larger than an electron, less deflection would occur.
18. $9.6 \times 10^{-7}$ N
19. The internal forces in an oxygen atom would be greater than those in a nitrogen atom because they have a greater number of charges. This would result in a stronger pull on the valence electrons and a smaller atomic radii.
20. The discovery of the neutron accounted for the unexpected mass of the nucleus and changed the idea that all atoms of the same element were identical.
21. The Rutherford nuclear model differed from the previous plum pudding model because nuclear model had a small dense positively charge core surrounded by mostly empty space. This model was based on evidence from the gold foil experiment in which alpha particles shot at a piece of gold foil mostly passed straight through (empty space) but some deflected backwards, having hit the nucleus.
22. Isotopes of a given element all have the same number of protons, but different numbers of neutrons and different masses.
23. 2.35%; 51
24. 16 u, 17 u, 18 u
25. Uncharged particles will not be deflected by a magnetic field.
26. Intensity. It is not possible to have 100% abundance of one isotope if the element has multiple stable isotopes.

27. 20.18 u; Ne

28. Copper

29. Chromium

30. Mass spectrometry can be used to determine the percent abundance of carbon-14 vs. carbon-12 in a sample.

31. The Rutherford model had electrons in orbit around the nucleus but did not account for Coulomb’s law which predicts that the electrons would fall into the nucleus. It could not explain why the orbiting electrons did not lose energy and produce a continuous spectrum of light.

32. Bohr explained spectral lines as produced when an electron fell from an excited orbit to a lower state orbit. This lines always produced specific wavelengths of light because the orbits had set values.

33. The wavelength of 698 nm is less energetic than the wavelength of 650 nm because the longer the wavelength the lower the frequency, $c=\lambda \nu$. The lower the frequency the lower the energy, $E=\hbar \nu$.

34. $2.85 \times 10^{-19}$ J
35. The \(2 \rightarrow 1\) transition is more energetic because electron orbits have a greater energy difference the closer they are to the nucleus.

36. Violet

37. 3

38. Atoms are bombarded with photons and the amount of energy required to remove each electron (ionization) is recorded.

39. Both electrons are in the same orbit and require the same amount of energy to ionize.

40. 3; 8

41. More than Bohr predicted.

42. 1400, the closer to the nucleus (smaller radii) the stronger the Coulombic forces and the greater the energy required to remove the electrons.

43. Lithium

44. De Broglie used evidence that electrons/matter behaves as both a particle and a wave to suggest that electron orbits did not lose energy because the electrons moved in a wave pattern around the nucleus.

45. In the Bohr model, these 8 electrons all exist on the same energy level/orbit. In the quantum model, the electrons are divided between two energy levels/subshells; 2 in the \(s\) subshell and 6 in the \(p\) subshell.

46. The quantum model predicts that boron has 3 energy levels; from left to right 2p, 2s, 1s.

47. 5 peaks. The 1s, 2s, 2p peaks will be of equal intensity with boron but of higher energy. The 3s and 3p peaks will be of equal intensity as the 1s and 2s peaks.

48. Chlorine

49. Left

50. 4

51. \(1s^22s^22p^5\)

52. \(1s^22s^22p^63s^23p^64s^23d^2; 1s^22s^22p^63s^23p^64s^2\)

53. \(1s^22s^22p^3; 1s^22s^22p^6\)
54. The Aufbau principle says electrons will fill the lowest energy subshells first. This is evidence by the intensity of the highest energy peaks on the PES spectra.

55. Elements in their excited state have the same number of electrons as the stable atom but one or more of these electrons is in a higher energy orbital that would be expected. An ion has actually lost or gained electrons.

56. 1s²2p¹
57. 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶5s²4d¹⁰5p⁶6s⁰5d¹

58. Bohr could not explain why the electrons stayed in stable orbits without losing energy. He assumed there was simply an exception to this law of physics.

59. The lowest energy state is the ground state, all others are excited.

60. Quantized means the value is a set multiple.

61. Emission spectra and absorption spectra are the inverse of one another. When an electron falls to a lower energy state it emits the same wavelength of light absorbed to excite the electron in the first place.

62. 102nm
63. 4.62x10⁻¹⁹ J; less
64. 3.00x10⁸ m/s

65. The photoelectric effect allows scientists to determine the binding energy of electrons by bombarding atoms with light to ionize the sample. In addition, it allows for solar electricity to be generated.

66. The intensity of the signal is determined by the number of electrons at a given energy level. The energy level is determined by the amount of energy required to remove those electrons, which is a product of the Coulombic force between those electrons and the nucleus.

67. This spectrum contradicts the Bohr model which predicts only two energy levels for carbon.

68. The energy levels of the peaks would be higher and the intensity of the lowest energy peak would be greater.

69. The smaller the radius between an electron and the nucleus the greater the Coulombic force. Greater force equates to greater energy to remove the electron.

70. Principal quantum number (n) represents the energy level of the shell. Angular quantum number (l) represents the shape of the orbital. Magnetic quantum number (m_l) represents the orientation of the orbital. Spin quantum number (m_s) represents the direction of the electron’s spin.

71. Ga would have one more peak than Fe. The 3d peak for Fe would have 3 times the intensity of its 4s peak. The 3d peak for Ga would have 5 times the intensity.

72. The Pauli exclusion principle says that each orbital can only hold 2 electrons. This is supported by the intensity of the peaks. For example, full s peaks have one-third the intensity of full p peaks because the s subshell has only one orbital, while the p subshell has 3.

73. The lowest energy peak represents the p subshell which contains only one electron compared to the two electrons in the s subshell.

74. Calcium’s highest and lowest energy peak both represent subshells containing two electrons.

75. 1s, 2p, 3p
76. The peak representing 2p contains 6 electrons all with the same binding energy, thus all degenerate.

77. 2s², 4s², 3d¹, 1s²

78. The binding energies of the 4s and 3d electrons are very similar.

79. Sc is unlikely to lose more than 3 electrons because after losing electrons from the 4s and 3d orbitals, the binding energy of the 3p orbital is 39.2, too high to be easily removed.

80. 1s²2s²2p⁶3s²3p⁶3d¹

81. Hund’s Rule means electrons will fill any empty orbitals first before pairing up to maximize stability and decrease electron-electron repulsion.

82. 0; 2

83. 1s²2s²2p⁶3s²3p⁶4s²3d⁶

84. 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰⁴p⁶⁵s²⁴d¹⁰⁵p⁵; 1s²2s²2p⁶3s²3p⁶⁴s²³d¹⁰⁴p⁶⁵s²⁴d¹⁰⁵p⁶

85. 1s²2s²2p⁶3s²3p⁶4s²3d⁶4p¹

86. Elements are arranged on the periodic table by increasing atomic number. Elements are groups by number of valence electrons and similar properties.

87. No, elements are identified by the number of protons they contain. Atoms can gain and lose electrons making it hard to determine the identity of the element.

88. Group 17

89. [Ne]3s²

90. [He]2s²2p⁶

91. Group 1, alkali metals

92. Aufbau principle because the ns orbitals and the (n-1)d orbitals often have similar binding energies.

93. Because they have a full subshell

94. Because they have the most repulsion between electrons without the benefit of stability from a completed shell or subshell.

95. [Ar]4s¹3d¹⁰

96. Periodic law says that properties of elements are a function of their atomic number.

97. Group 18 is the noble gases. These elements are inert due to the stability provided by the full shell of electrons.

98. [Ar]4s²3d¹⁰⁴p⁴

99. [Ar]4s²3d⁹⁴p⁶

100. Transition metal

101. Iron has a half full sub-shell

102. Yb and No – full subshell, Eu and Am – half full

103. [Ar]4s¹3d⁵

104. Chlorine’s configuration is not an exception because the valence electrons in the p orbital are much lower than those in the d or s orbitals.

105. Atomic radius increases down a group because energy level shell are added.

106. We would expect the radius to be larger because of the added energy shell. This does not occur as expected because of the addition of d orbital electrons.

107. Fluorine would have a smaller radius, this can be determined from the higher energy values for F on the PES spectrum. The higher energy values are caused by an
increase in Coulombic force due in part to smaller radial distance between electrons and the nucleus.

108. Greater effective nuclear charge is responsible for the difference in radii. These elements have the same shielding constant because they have the same number of inner electrons.

109. Fluorine has more electrons in its p orbital.

110. Magnesium ions are smaller than magnesium atoms because they experience less electron-electron repulsion.

111. Si < Ge < Ga

112. Ar < S < P < P\(^3^–\)

113. Electronegativity is the pull of electrons toward the nucleus of an atom. Noble gases have no measurable electronegativity values because they have a full shell of electrons preventing the nucleus of the atom from attracting additional electrons.

114. Electronegativity graph

115. P > Al > Mg > Ca

116. F – less shielding; O – much higher nuclear charge; O – higher nuclear charge – same shielding; Cl – much higher nuclear charge even though the nucleus is shielded more.

117. The energy required to remove an electron from an atom. It is influenced by:

   Nuclear charge – the higher it is, the higher the IE; Shielding – the higher it is, the lower the IE; Orbital effects – full or half-full orbitals render stability to the atom and increase the IE.

118. The shielding is constant but Mn has the extra proton in the nucleus thereby increasing the nuclear charge increasing the IE

119. 4s; 3d; 3p

120. Sc

121. The 4s orbital because it has the lowest binding energy.

122. 2+ and 3+ ions would be likely because 2 electrons can be lost from the 4s orbital and 1 from the d subshell. A

123. 4+ ion would likely not occur because the energy required to remove an electron from the p subshell is much higher.

124. Metallic elements have low ionization energies because they have smaller effective nuclear charges resulting in less Coulombic force between the valence electrons and the nucleus and thus lose electrons easily.

125. Bromine has a higher effective nuclear charge and higher electronegativity causing it to pull electrons toward it, while these values for copper are lower, allowing it to lose electrons and conduct electricity.

126. Metallic character graph
Metallic character

Atomic Number

127. Because the next period adds d orbitals.
128. $4s^2$ then $3d^5$; $4s^2$ then $3d^{10}$; $4s^2$ then $4s^2$; $5p^2$ and $5s^2$; $5s^2$; $5s^1$
129. It also has 6 full energy levels of shielding between the nucleus and valence electrons. All other elements – such as Ra or U have more protons than Fr for its period.
130. Lithium’s valence electron is in the 2s orbital so the 1s orbital acts to shield the nuclear charge thereby lessening the Coulombic attractions and allowing the atom to expand. With helium, there is no shielding between the valence electrons and the nucleus so the attractions are strong and the radii is small. All of the elements in this period are shielded by the first two energy levels so it is the nuclear charge that factors. As we move across the period, the atomic number and nuclear charge increase thereby increasing the attractive forces and shrinking the radii.
131. Fluorine has a smaller atomic radii because it has a higher effective nuclear charge than oxygen.
132. Core – they cancel out the nuclear charge so the valence electrons feel less Coulombic attraction and therefore expand outward.
133. We cannot know the precise location of an electron so we cannot know the precise boundary of an atom.
134. Less than, both ions are isoelectronic with neon and thereby see the same amount of shielding. The magnesium ion however has a much higher nuclear charge and thereby makes the ion smaller.
135. Ca > Br$^-$ > Br > Ca$^{2+}$
136. Cl$^-$ > Cl > Ar
137. It decreases due to the extra shielding each element down the group has.
138. The extra shielding diminishes the nuclear charge felt by chlorine’s valence electrons compared to fluorine’s valence electrons.
139. N > S > C > Si
140. Ionization energy decreases with increasing atomic radii, the larger the element the easier it is to remove a valence electron. Ionization energy increases with increasing electronegativity, the greater the effective nuclear charge and smaller the atom the greater the attraction between electrons and the nucleus.
141. Noble gases have the highest ionization energies of their periods because they are the smallest element in their period and experience the greatest effective nuclear charge.
142. K < Na < Sc
143. For 1st IE, calcium’s will be greater due to the higher nuclear charge for a given level of shielding. For 2nd IE, potassium’s will be greater due to its’ being isoelectronic with argon thereby losing a level of shielding and enjoying the stability of a full valence shell.
144. Chlorine has a greater electronegativity due to it higher effective nuclear charge and thus attracts electrons to form anions, while aluminum has lower 1st, 2nd, and 3rd ionization energies due to its larger size and lower effective nuclear charge.
145. Al loses its outermost 3p and 3s electrons. Sc loses its outermost 4s then 3d electrons.
146. Mg → Mg⁺ + e⁻; Mg⁺ → Mg²⁺ + e⁻; Mg²⁺ → Mg³⁺ + e⁻
147. Lead has a larger atomic radius due to its number of energy levels (6 to carbon’s 2).
   The larger the radius the smaller the Coulombic attraction between the valence electrons and the nucleus, thus lead’s valence electrons experience lower binding energy are more easily lost (metallic) than carbon’s.
148. Fr>Sr>K>Na>Ga
149. Silver is less metallic than copper. Silver forms 1+ ions while copper forms 2+ or 4+ ions – losing more electrons = more metallic. Note that this does not follow the general trend.
150. Atoms within the same group have the same valence shell configuration whereas atoms in the same period all have different valence shell configurations.
151. Lead is lose 6p² and then 6s²
152. 2+ and 4+; 1+ and 3+

1. Tungsten
   a. 74 protons; 106 neutrons; 108 neutrons; 109 neutrons; 110 neutrons; 112 neutrons.
   b. The percent abundance can be obtained using a mass spectrometer to ionize samples of tungsten and determine the relative abundance of each mass/charge ratio in the sample.
   c. Mass spec
   d. [Xe]6s⁴4f¹⁵d⁵
2. Unknown mass spec.
a. Mass spec

b. 35.5, Chlorine
c. 17 protons; 18 neutrons; 20 neutrons
d. The discovery of isotopes altered our understanding of the mass of atoms because we now know that not all atoms of a given element have the same mass.

3. Unknown elements
   a. Bromine or Iodine
   b. Niobium or Rhodium
   c. Zinc
   d. Lithium
   e. Yttrium
   f. Carbon or Oxygen
   g. Radium

4. Chalcogens
   a. Po, it has the most shielding so loses electrons easily.
   b. Oxygen v. Fluorine
      i. F, higher nuclear charge for given amount of shielding.
      ii. F, higher nuclear charge, stronger attractions and smaller radius

5. Ionization energies
   a. The shielding is constant across the period but nuclear charge increases thereby increasing the IE
   b. Exceptions to the trend (lowering IE) occur due to stability provided by full shell and subshells across the period.
   c. It will be less than that of both Li and Ne due to the increased shielding of the nucleus by the now full 2nd energy level.
   d. B^{2+} because it will lose its 2s electrons. Be^{3+} because it will lose its 2p electron and its 2s electrons.
6. Lithium & Helium
   a. PES spectra
      i. \(2s^1\)
      ii. \(1s^2\)
      iii. \(1s^2\)
   b. Lithium has the higher nuclear charge and the shielding is the same for both elements 1s orbital.
   c. Ionization energies
      i. Helium due to less shielding and a full valence shell
      ii. Lithium as it is now isoelectronic with helium with a full valence shell but also has the greater nuclear charge for the same amount of shielding.
   d. These spectra would support the Bohr model because they have 1 peak per energy level \((n)\) as predicted by Bohr.

7. Sulfur
   a. The peak with binding energy of 1.0
   b. \(1s^22s^22p^63s^23p^4\)
   c. Higher due to the higher nuclear charge due to the extra proton chlorine has in its nucleus.
   d. The sulfide ion would have a larger radius than the sulfur atom due to increase electron-electron repulsion.
   e. This spectrum does not support the Bohr model. The Bohr model would predict only 3 peaks, one for each energy level.