Chapter 8 Reading Guide **Name:** ____________________
AP Chemistry 2016-2017 **Date:** **Period:**

Understanding the arrangement of the elements in the periodic table is key foundational knowledge in chemistry. The relationships and patterns observed explain trends of the elements. The emphasis in AP Chemistry is the explanation of the concepts, not memorizing trends. For example, the size of the atoms in Group I decreases as you go down a column. This is a trend. However, the explanation of the trend requires students to know how to apply Coulomb's force law to charged particles and effective nuclear charge. This material is found in the AP Chemistry curriculum in Big Ideas 1, 2, 3, and 5.

8.1 Nerve Signal Transmission

1. How are ion pumps in cells able to differentiate between two group 1 metal ions, sodium and potassium?

2. Explain what a periodic property is and give an example.

8.2 The Development of the Periodic Table

3. How did Mendeleev arrange the elements in the periodic table, and why has his arrangement survived to the present day?

4. What elements did Mendeleev predict would be discovered?

5. How did Mosely modify Mendeleev's organization of the elements?

8.3 Electron Configurations: How Electrons Occupy Orbitals

6. What is meant by electron configuration? What does it describe?

7. Define the **ground state** of an atom.

8. What is sublevel-energy splitting?

9. Draw an orbital diagram of a hydrogen atom.

10. What are the two possible spin states of an atom?

11. Explain the **Pauli exclusion principle**.

12. Describe **<u>three</u>** main ideas from Coulomb's laws about attractions and repulsions.
    a.

    b.

    c.

13. What is **shielding** in an atom, and what does it help explain about atomic properties?

14. What is effective nuclear charge? What does it help explain about atomic properties?

15. What is the difference between shielding and effective nuclear charge? How do these two concepts together help explain atomic properties?

16. Why does a 2*s* orbital experience greater nuclear charge than 2*p* orbitals?

17. Define:
    a. The Aufbau principle

    b. Hund's rule

18. Summarize the rules of orbital filling.

19. Write electron configurations and make orbital drawings for the following atoms:
    a. Carbon

    b. Oxygen

    c. Sodium

8.4 Electron Configurations, Valence Electrons, and the Periodic Table

20. How are **core** and **valence electrons** different? Use an example in your explanation.

21. In a ground state atom, are *d* and *f* orbital electrons considered valence electrons? Why or why not? Give an example.

22. Describe the outer electron configuration of:
    a. Alkali metals
    b. Alkaline earth metals
    c. Halogens
    d. Noble gases

23. Where are the inner transition metals located on the periodic table? Why are they called inner?

24. Sketch a periodic table. Identify the location of atoms that are filling their last electrons in *s*, *p*, *d*, and *f* orbitals.

25. Why is the *s* block filled only in two columns of the periodic table?

26. How can you look at the periodic table and determine the number of valence electrons in any ground-state atom?

27. How does the principle quantum number (energy level) of *d* orbitals compare to its row in the periodic table? Of *f* orbitals?

Note: on the AP exam, you will not be expected to have memorized exceptions to the Aufbau Principle. However, upon being told that a given element is an exception, you should be able to provide reasoning for why the exception occurs

8.5 The Explanatory Power of the Quantum-Mechanical Model

28. For group 1 elements, explain how the electron configuration corresponds to the properties (such as reactivity) of elements in the group.

29. Why are noble gases the most chemically stable? Why is Xe less stable than Ne?

30. Which group of metals is the most reactive? Explain why.

31. Which group of non-metals is the most reactive? Explain why.

8.6 Periodic Trends in the Size of Atoms and Effective Nuclear Charge

32. Define **atomic radius**, including a drawing in your answer.

33. Describe and explain the trend in atomic radius as you go down a group on the periodic table.

34. Describe and explain the trend in atomic radius as you go across a period on the periodic table.

35. How is the effective nuclear charge, Z_{eff}, of an atom determined?

36. How do the radii of transition metals change as you move across a period in the periodic table? Explain the trend.

8.7 Ions: Electron Configurations, Magnetic Properties, Ionic Radii, and Ionization Energy

37. Explain how to determine the number of electrons in an anion. Use O^{2-} as an example.

38. Explain how to determine the number of electrons in a cation. Use Na^{+} as an example.

39. Explain why the last electron orbital to be occupied in the neutral atom is not always the first electron to be removed when the atom forms a cation. Provide an example.

40. Define paramagnetic and diamagnetic. Explain how each can be predicted, using examples.

41. How does the atomic radius change when a cation is formed? Explain why.

42. How does the atomic radius change when an anion is formed? Explain why.

43. Define **isoelectronic**, and give an example of three isoelectronic species in your answer.

44. Define **ionization energy**. Write a chemical equation showing the process of ionization energy.

45. Explain the difference between IE_1, IE_2, and IE_3 using lithium as an example.

46. Describe the trend in ionization energy in the period from Li to Ne.

47. Explain the trend in ionization between Be and B. Explain why it occurs.

48. Explain the trend in ionization between N and O. Explain why it occurs.

49. When comparing successive ionization energies of elements in the periodic table, when would a large increase in the ionization energy of an atom be expected to be observed? Explain.

8.8 Electron Affinities and Metallic Character

50. What is **electron affinity**?

51. Write a chemical equation representing the process of electron affinity.

52. Explain the electron affinity trends in the main group elements:
    a. Across a period

    b. Down a group

53. What causes these trends in electron affinity?

54. Define the following terms:
    a. Conductivity

    b. Malleability

    c. Ductility

55. Identify typical trends of metallic character:
    a. Down a group (use group 5A as an example)

    b. Across a period (use period 4)

8.9 Some Examples of Periodic Chemical Behavior: The Alkali Metals, the Halogens, and the Noble Gases

56. Explain the trend in reactivity within the group of alkali metals. Explain why this trend occurs.

57. Write a chemical equation showing how an alkali metal reacts with water. You pick the alkali metal.

58. What does it mean for an atom to be reduced? Why are the halogens so easily reduced?

59. Which element is most likely to react with the heavy noble gases to form stable compounds?

60. How do halogens react with metals? Include a chemical equation in your answer.

61. Give an example of a metal halide.

Self-Assessment Answers

1. ________	4. ________	7. ________	10. ________	13. ________
2. ________	5. ________	8. ________	11. ________	14. ________
3. ________	6. ________	9. ________	12. ________	15. ________