Structures and Properties of Matter Review

1. In what ways were Dalton's and Thomson's atomic model similar? In what ways were they different?

2. In what ways were Thomson's and Rutherford's models of the atom similar? In what ways were they different?

3. What is common to all atomic models, from Dalton's to Bohr's?

4. Describe how the motion of an electron in Bohr's model of the atom is different from the motion of an electron in the quantum mechanical model of the atom.

5. Explain how periodic trends in ionization energy and electron affinity explain why atoms of elements in Group 1 and Group 2 tend to bond with other elements by forming positive ions in ionic compounds.

6. A nitrogen atom has a total of three electrons in its 2p orbitals. Are any of these electrons paired? Explain how you know.

7. What information do the quantum numbers provide about an orbital?

8. What two things does the principal quantum number (n) describe about an orbital and the electron in it?

9. What is the Pauli exclusion principle and how is it related to the spin quantum number (m_s)?

10. What is Hund's rule and its significance?

11. Why does it make sense that electron configurations represent atoms in their ground state and not their many excited states?

12. What are the general trends for atomic radius down a group and across a period in the periodic table?

13. Without consulting a periodic table, determine to which period, group, and block the element with the electron configuration $[Xe]6s^2 5d^{10}$ would belong. Check your answer by using your periodic table.

14. Write the complete electron configuration for the element in Period 5 and Group 15.

15. List the 3s, 5s, 2p, 4p, 3d, and 5d orbitals in order, from the lowest energy to the highest energy.

16. Which element is defined by:
    a. $[Ar]4s^2 3d^8$?
    b. $[Ne]2s^2 2p^1$?
    c. $[Kr]5s^2 4d^{10} 5p^4$?

17. Referring only to a periodic table, arrange the following sets of elements in order of increasing first ionization energy. Briefly explain your reasoning.
    a. Na, Si, Ar
    b. Mg, Ca, Ba
    c. He, Li, Be

18. Consider the electron configuration: $1s^2 2s^2 2p^4$.
    a. Assume the electron configuration represents a neutral atom in its ground state. What element does it represent? Explain how you know.
    b. What information does this electron configuration notation provide?
    c. What information does this electron configuration notation not provide?

19. Electrons are said to be "delocalized" in metallic bonds.
    a. In your own words, describe what this means, and how it is different from the locations of valence electrons in ionic and covalent bonds.
    b. How does the "delocalization" of electrons in metallic solids explain "malleability", the property unique to all metals?

20. Explain how single bonds form, based on quantum mechanical concepts.

21. What experimentally observed property of methane makes it necessary to invoke the concept of hybridization to explain the structure of methane?

22. List five possible shapes of hybrid orbitals.

23. Describe two ways in which a non-polar molecule can temporarily become a dipole.

24. Explain why symmetrical molecules are non-polar and asymmetrical molecules can be polar.

25. Classify the following bonds as mostly ionic, polar covalent, or mostly covalent by looking at the location of the elements on the periodic table. Check your classifications by calculating ^EN for each.
    a. Li-Cl
    b. S-S
    c. C-N
    d. Na-O

26. Use the periodic table to help you write the condensed electron configuration for the following elements and their most likely ion (if there is one):
    a. Lithium
    b. Argon
    c. Chlorine
    d. Phosphorus

27. Use VSEPR theory to identify the electron group arrangement (VSEPR shape), the molecular shape, and the bond angle of the following molecules whose central atoms have:
    a. 4 bonding pairs and 1 lone pair
    b. 6 bonding pairs and 0 lone pairs
    c. 3 bonding pairs and 2 lone pairs
    d. 3 bonding pairs and 0 lone pairs
    e. 2 bonding pairs and 2 lone pairs
    f. 4 bonding pairs and 2 lone pairs

28. Determine whether each of the following compounds will be polar or non-polar.
    a. CO_2
    b. H_2S
    c. SiO_2
    d. PCl_3

29. In the 1950s, the reaction of hydrazine (N_2H_4) with chlorine trifluoride (ClF_3) was used as a rocket fuel.
    a. Draw the Lewis structures for hydrazine and chlorine trifluoride.
    b. Identify the hybrid orbitals used in each one.
    c. Identify the molecular shape and polarity of chlorine trifluoride.

30. Use VSEPR theory and Lewis structures to predict the number on bonding pairs and lone pairs around the central atom so that you can identify the VSEPR shape, molecular shape, and bond angle for the following molecules and ions:
    a. XeF_2
    b. BCl_4^-
    c. SF_5^+

31. Which compound in each of the following pairs has the higher boiling point? Explain your choice in each case.
    a. NH_3 or PH_3
    b. C_2H_6 or C_4H_{10}
    c. $SeCl_4$ or $SiCl_4$

32. For each of the following elements or compounds, predict which would have the higher boiling point and explain how you made your choice.
    a. O_2 or N_2
    b. Ethanol or methoxymethane
    c. Heptane or 2,4-dimethylpentane