Date: **Name:**

SCH4U	Worksheet
Structure and Properties of Matter	

Worksheet

Learning Check

1. What evidence invalidated Dalton's atomic theory that atoms are indivisible?

2. If Thomson's model of the atom had been valid, what would Rutherford and his team have observed in the gold foil experiment, and why?

3. State one way the discovery of radioactive elements furthered the development of atomic theory.

4. Draw two labelled diagrams with captions to summarize Thomson's and Rutherford's atomic models.

5. Dalton enjoyed lawn-bowling, and plum pudding was a popular dessert in England long before Thomson was even born. What role do you think that personal interests and cultural history play in the kinds of models that scientists use to communicate their ideas? Use another example to explain your answer.

6. Explain why Rutherford and his students expected to observe only minor deflections of alpha particles, based on Thomson's model.

Learning Check

7. What is the key difference between the orbitals of a hydrogen atom and the orbitals of multi-electron atoms?

8. Write the following sets of orbitals in order from lowest to highest energy.

a. $2s$, $2p$, $3s$, $3p$

b. $3p$, $3d$, $4s$, $4p$

c. $4d$, $4f$, $5s$, $5p$, $5d$, $5f$, $6s$, $6p$

9. What does it mean to say that an orbital is "full"?

10. How many electrons can occupy all possible orbitals with $n = 1$ and $n = 2$? Show two ways to arrive at the answer.

11. The orbital diagram for helium is given in Figure 3.18. Would drawing two arrows pointing in the same direction also be a correct orbital diagram? Explain your answer.

12. Draw an orbital diagram to represent each of the following:

a. an "unoccupied" orbital

b. an orbital with a single electron

c. a filled orbital

Learning Check

13. Why do electrons fill the $5s$ orbital before the $4d$?

14. Write full and condensed electron configurations and orbital diagrams for boron and neon.

15. Write the complete electron configuration for sulfur.

16. Draw a partial orbital diagram for silicon in the ground state.

17. Identify the elements with these condensed electron configurations, and write the complete forms.

a. $[Ne]3s^1$ **b.** $[Ar]4s^2 3d^3$

18. Identify the following element, and write a condensed electron configuration for it.
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

Learning Check Answers

SCH4U	**Worksheet**
Structure and Properties of Matter	

1. Thomson's discovery of the electron in 1897 invalidated Dalton's atomic theory.

2. Alpha particles would have passed straight through the foil with minimal or no deflection from encounters or collisions with nearby electrons. There would be no deflection caused by the positive charge because Thomson's model postulates a uniform, positive charge spread throughout the atom.

3. Some radioactive elements emit positively charged alpha particles. Rutherford studied them and then used the alpha particles to bombard thin foils including gold foils. This led to the model in which all of the positive charge and most of the atom's mass were confined to a very small region at the centre of the atom, which Rutherford called the nucleus.

4. Diagrams should be based on Figures 3.3 and 3.6. Both models are spherical and include electrons and the positive charge. In Thomson's model, the positive charge is spread throughout the sphere and electrons are embedded in the sphere like raisins in a muffin. In Rutherford's model, the positive charge is found in a tiny, extremely dense nucleus and the electrons orbit the nucleus like planets.

5. Scientists tend to name their models, or other discoveries, after something that is common to their own everyday lives. Rutherford's model is sometimes called the planetary model.

6. In Thomson's model, negative charges were scattered evenly throughout a large positively charged mass. The alpha particles were highly energetic and would not be expected to be deflected very much by such atoms.

7. In a hydrogen atom, orbital energy depends only on n. For example, electrons in $2s$ and $2p$ have the same energy. In multi-electron atoms, orbitals in different sublevels have different energies associated with them, even if they have the same value of n. For example, $2s$ and $2p$ are associated with different energies.

8. **a.** $2s$, $2p$, $3s$, $3p$
b. $3p$, $4s$, $3d$, $4p$
c. $5s$, $4d$, $5p$, $6s$, $4f$, $5d$, $6p$, $5f$

9. An orbital is "full" when it contains two electrons.

10. Method one: There are five possible orbitals for $n = 1$ and $n = 2$: one $1s$ orbital, one $2s$ orbital, and three $2p$ orbitals. Each of these can contain a maximum of two electrons. Therefore, 10 electrons can occupy all possible orbitals with $n = 1$ and $n = 2$. Method two: Using the formula $2n^2$, $n = 1$ can contain two electrons and $n = 2$ can contain eight electrons, for a total of 10.

11. No. Two arrows pointing in the same directions would indicate that two electrons in the same orbital have the same spin quantum number. This violates the statement made in the Pauli exclusion principle that no two electrons can have the same four quantum numbers.

12. **a.** $\boxed{}$ **b.** $\boxed{\uparrow}$ **c.** $\boxed{\uparrow\downarrow}$

13. Orbitals fill in order of increasing energy. At energy levels above $n = 3$, the different sublevels overlap. As a result, the $5s$ orbital has a lower energy than the $4d$ orbitals.

14. boron: $1s^2 2s^2 2p^1$; $[He]2s^2 2p^1$
neon: $1s^2 2s^2 2p^6$; $[He]2s^2 2p^6$

15. $1s^2 2s^2 2p^6 3s^2 3p^4$

16. $\underset{3s}{\boxed{\uparrow\downarrow}}\ \underset{3p}{\boxed{\uparrow}\boxed{\uparrow}\boxed{}}$

17. **a.** sodium: $1s^2 2s^2 2p^6 3s^1$
b. vanadium: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

18. titanium: $[Ar]4s^2 3d^2$

SCH4U	Worksheet
Structure and Properties of Matter	

Practice Problems

1. For the quantum number $n = 3$, what values of l are allowed, what values of m_l are possible, and how many orbitals are there?
2. If $n = 5$ and $l = 2$, what orbital type is this, what are the possible values for m_l, and how many orbitals are there?
3. What are the n, l, and possible m_l values for the following orbital types?
   a. $2s$ **c.** $5d$
   b. $3p$ **d.** $4f$
4. What orbital type can be described by the following sets of quantum numbers?
   a. $n = 2, l = 0, m_l = 0$
   b. $n = 5, l = 3, m_l = -2$
5. How many orbitals are associated with each of the following types?
   a. $1s$ **c.** $4f$
   b. $5f$ **d.** $2p$
6. What sets of quantum numbers are possible for a $4d$ orbital? List them.
7. What is one possible value for the missing number in each of the following sets?
   a. $n = 3, l = 1, m_l = ?$
   b. $n = 2, l = ?, m_l = -3$
8. Write two possible sets of quantum numbers for a $6p$ orbital.
9. The following sets of quantum numbers are not allowed. Identify the problem and change one number to give an allowed set.
   a. $n = 1, l = 2, m_l = -2$
   b. $n = 4, l = 1, m_l = -2$
10. Label each of the following sets of quantum numbers as *allowed* or *not allowed*. Identify the problem for each of the *not allowed* sets.
    a. $n = 3, l = 2, m_l = 0$
    b. $n = 1, l = 1, m_l = -1$
    c. $n = 0, l = 0, m_l = 0$
    d. $n = 5, l = 1, m_l = 3$

Practice Problems

11. Write complete and condensed electron configurations for yttrium, Y.
12. Write complete and condensed electron configurations for lead, Pb.
13. What elements have the valence electron configuration that is given by ns^2?
14. What elements have the valence electron configuration that is given by $ns^2(n-1)d^3$?

15. What are the two exceptions to the guidelines for filling orbitals in Period 4? Draw the partial orbital diagrams you would expect for them, based on the aufbau principle. Then draw partial orbital diagrams that represent their actual electron configurations. Finally, explain why the discrepancy arises.

16. The condensed electron configuration for strontium is $[Kr]5s^2$. Without using a periodic table, identify the group number to which strontium belongs. Show your reasoning.

17. Identify the following elements and write condensed electron configurations for atoms of each element.

a. The *d*-block element in Period 4 with 10 valence electrons.

b. The element in Period 6 with 3 valence electrons.

18. The condensed electron configuration for titanium is $[Ar]4s^2 3d^2$. Without a periodic table, identify the period number to which titanium belongs. Show your reasoning.

19. The condensed electron configuration for arsenic is $[Ar]4s^2 3d^{10} 4p^3$. Without using a periodic table, identify the orbital block in the periodic table to which arsenic belongs. Show your reasoning.

20. Without a periodic table, and based on the condensed electron configuration given below, identify the group number, period number, and orbital block to which each of the following elements belongs. Show your reasoning.

a. francium, $[Rn]7s^1$

b. tungsten, $[Xe]6s^2 4f^{14} 5d^4$

c. antimony, $[Kr]5s^2 4d^{10} 5p^3$

1. l can be 0, 1, or 2; for $l = 0$, $m_l = 0$; for $l = 1$, m_l can be −1, 0, or +1; for $l = 2$, m_l can be −2, −1, 0, 1, or 2; There are 9 orbitals.

2. 5*d*; m_l can be −2, −1, 0, +1, or +2; There are five orbitals.

3. a. $n = 2, l = 0, m_l = 0$

b. $n = 3, l = 1, m_l = -1, 0, \text{ or } +1$

c. $n = 5, l = 2, m_l = -2, -1, 0, +1, \text{ or } +2$

d. $n = 4, l = 3, m_l = -3, -2, -1, 0, +1, +2, \text{ or } +3$

4. a. 2*s*

b. 5*f*

5. a. 1

b. 7

c. 7

d. 3

6. $n = 4, l = 2, m_l = -2$; $n = 4, l = 2, m_l = -1$; $n = 4, l = 2, m_l = 0$; $n = 4, l = 2, m_l = +1$; $n = 4, l = 2, m_l = +2$

7. a. $m_l = -1$

b. $l = 3$

8. $n = 6, l = 1, m_l = -1$; $n = 6, l = 1, m_l = 0$ (many more)

9. Sample answers:

a. $l = 2$ does not exist with $n = 1$; change $n = 1$ to $n = 3$

b. $m_l = -2$ does not exist with $l = 1$; change $l = 1$ to $l = 2$

10. a. Allowed

b. Not allowed; $l = 1$ does not exist with $n = 1$

c. Not allowed; $n = 0$ does not exist

d. Not allowed; m_l can be only −1, 0, or +1 when $l = 1$

11. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^1$; $[Kr]5s^2 4d1$

12. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$; $[Xe]6s^2 4f^{14} 5d^{10} 6p^2$

13. Group 2 elements

14. Group 5; V, Nb, Ta, Db

15. The two exceptions are copper and chromium. The expected diagram for chromium would have a filled 4*s* orbital and four 3*d* orbitals containing one electron each (the fifth 3*d* orbital would be empty.) The expected diagram for copper would have a filled 4*s* orbital, four filled 3*d* orbitals, and one 3*d* orbital with one electron. The actual diagram for chromium has one electron in its 4*s* orbital and one electron in each of its five 3*d* orbitals. The actual diagram for copper has one electron in its 4*s* orbital and five filled 3*d* orbitals. The discrepancy arises because the predicted electron configurations have a slightly higher energy than the actual electron configurations. When the 3*d* orbitals are either all half-filled or all completely filled, the configuration becomes stable.

16. The configuration for the valence electrons, s^2, indicates that strontium is in Group 2.

17. a. nickel: $[Ar]4s^2 3d^8$

b. lanthanum, La: $[Xe]6s^2 5d^1$

18. The value of 4 in $4s^2$ indicates that titanium is in Period 4.

19. The configuration of the valence electrons shows full *s* and *d* orbitals, but a half-full *p* orbital, so arsenic must be in the *p* block.

20. a. The electron configuration of the valence electron is s^1, so francium must be in Group 1 and belong to the *s* block. Because of the 7 in $7s^1$, francium must be in Period 7.

b. The electron configuration of the valence electrons is $s^2 f^{14} d^4$, so tungsten must be in Group 6 (2 + 4) and belong to the *d* block. Because of the 6 in $6s^2$, tungsten must be in Period 6.

c. The electron configuration of the valence electrons is $s^2 d^{10} p^3$, so antimony must be in Group 15 (2 + 10 + 3) and belong to the *p* block. Because of the 5 in $5s^2$, antimony must be in Period 5.