Exercise 4.2
Quantum Number Rules

1. Does each of the following set of quantum numbers describe a possible atomic orbital? If so, give the label for this orbital. If not, explain why an electron with that set of quantum numbers isn't possible.

(a) $n = 4,\ l = 3,\ m_l = -3,\ m_s = ½$

(b) $n = 1,\ l = 0,\ m_l = 0,\ m_s = -1$

(c) $n = 0,\ l = 0,\ m_l = 0,\ m_s = -½$

(d) $n = 2,\ l = 3,\ m_l = 0,\ m_s = ½$

(e) $n = 3,\ l = -2,\ m_l = 3,\ m_s = -½$

2. We introduce quantum numbers and quantum number rules in CHEM 1000 because they dictate the maximum possible number of electrons in each type of orbital. This is a direct consequence of the Pauli exclusion principle.

(a) State the Pauli exclusion principle.

(b) What is the maximum number of electrons that can be in a 2p orbital in a single atom?

(c) What is the maximum number of electrons that can be in a 2p subshell in a single atom?

(d) Briefly explain how the Pauli exclusion principle helped you answer parts (b) and (c).

3. What is the maximum number of electrons (in a single atom) that can be associated with each of the following combinations of quantum numbers?

(a) $n = 2$

(b) $n = 3,\ l = 2$

(c) $n = 2,\ l = 0,\ m_l = 0$

(d) $n = 6,\ l = 3,\ m_l = -3,\ m_s = -½$