Answers to 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) yes — 4f

(b) no — m_s must be either +½ or -½

(c) no — n cannot be 0

(d) no — l cannot be greater than n

(e) no — l cannot be negative
also, since m_l cannot be greater than l and l cannot be greater than or equal to n, m_l cannot be equal to n

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) No two electrons in the same atom can have the same set of quantum numbers (the same "quantum state").

(b) 2

(c) 6

(d) **explaining part (b)**
A single 2p orbital will have $n = 2$, $l = 1$ and one legal value of m_l. For all electrons in that orbital to have different sets of quantum numbers, they must each have a different value for m_s. Since there are only two legal values for m_s for an electron, there can only be two electrons in that 2p orbital.
explaining part (c)
The 2p subshell is the set of all 2p orbitals in an atom. They will all have $n = 2$ and $l = 1$ but there are three different legal values for m_l (-1, 0 and +1), indicating that there are three different 2p orbitals in the subshell (one orbital with each m_l value). As explained above, each of those three orbitals can contain two electrons (m_s = +½ and m_s = -½) so there can be, at most, six 2p electrons in an atom.

3.

(a) 8

(b) 10

(c) 2

(d) 1