1. Wavelengths of the hydrogen spectrum are given by $1/\lambda = R_H(1/n_f^2 - 1/n_i^2)$, where the Rydberg constant is $R_H = 1.097\,373\,2 \times 10^7$ m^{-1}. Thus, with $n_f = 3$ and $n_i = 5$,

$$\frac{1}{\lambda} = 1.097\,373\,2 \times 10^7 \text{ m}^{-1} \left(\frac{1}{3^2} - \frac{1}{5^2} \right) = 7.80 \times 10^5 \text{ m}^{-1},$$

and $\lambda = 1/(7.80 \times 10^5 \text{ m}^{-1}) = 1.28 \times 10^{-6}$ m, so (a) is seen to be the correct choice.

3. With a principal quantum number of $n = 3$, there are 3 possible values of the orbital quantum number, $\ell = 0,\ 1,\ 2$. There are a total of $2(2\ell + 1)$ possible quantum states for each value of ℓ; $2\ell + 1$ possible values of the orbital magnetic quantum number m_ℓ, and 2 possible spin orientations $(m_s = \pm\frac{1}{2})$ for each value of m_ℓ. Thus, there are 10 $3d$ states (having $n = 3,\ \ell = 2$), 6 $3p$ states (with $n = 3,\ \ell = 1$), and 2 $3s$ states (with $n = 3,\ \ell = 0$), giving a grand total of $10 + 6 + 2 = 18$ $n = 3$ states and the correct choice is (e).

4. There are 6 distinct possible downward transitions with 4 energy levels. These transitions are: $4 \rightarrow 1,\ 4 \rightarrow 2,\ 4 \rightarrow 3,\ 3 \rightarrow 1,\ 3 \rightarrow 2$, and $2 \rightarrow 1$. Thus, assuming that each transition has a unique photon energy, $E_{photon} = |\Delta E| = E_i - E_f$, associated with it, there are 6 different wavelengths $\lambda = hc/E_{photon}$ the atom could emit and (e) is the correct choice.

5. The structure of the periodic table is the result of the Pauli exclusion principle, which states that no two electrons in an atom can ever have the same set of values for the set of quantum numbers $n,\ \ell,\ m_\ell$, and m_s. This principle is best summarized by choice (c).

6. All states associated with $\ell = 2$ are referred to as d states. Thus, all 10 possible quantum states having $n = 3,\ \ell = 2$ are called $3d$ states (see Question 3 above), and the correct answer is choice (c).

7. Of the electron configurations listed, (b) and (e) are not allowed. Choice (b) is not possible because the Pauli exclusion principle limits the number of electrons in any p subshell to a maximum of 6. Choice (e) is impossible because the selection rules of quantum mechanics limit the maximum value of ℓ to $n - 1$. Thus, a $2d$ state $(n = 2,\ \ell = 2)$ cannot exist.

2. Neon signs do not emit a continuous spectrum. They emit many discrete wavelengths, as could be determined by observing the light from the sign through a spectrometer. However, they do not emit all wavelengths. The specific wavelengths and intensities account for the color of the sign.

8. The de Broglie wavelength of macroscopic objects such as a baseball moving with a typical speed such as 30 m/s is very small and impossible to measure. That is, $\lambda = h/mv$ is a very small number for macroscopic objects. We are not able to observe diffraction effects because the wavelength is much smaller than any aperture through which the object could pass.

28.1   (a)   The wavelengths in the Lyman series of hydrogen are given by $1/\lambda = R_H(1 - 1/n^2)$, where $n = 2, 3, 4, \ldots$, and the Rydberg constant is $R_H = 1.097\ 373\ 2 \times 10^7$ m^{-1}. This can also be written as $\lambda = (1/R_H)(n^2/n^2 - 1)$ so the first three wavelengths in this series are

$$\lambda_1 = \frac{1}{1.097\ 373\ 2 \times 10^7\ \text{m}^{-1}}\left(\frac{2^2}{2^2 - 1}\right) = 1.215 \times 10^{-7}\ \text{m} = \boxed{121.5\ \text{nm}}$$

$$\lambda_2 = \frac{1}{1.097\ 373\ 2 \times 10^7\ \text{m}^{-1}}\left(\frac{3^2}{3^2 - 1}\right) = 1.025 \times 10^{-7}\ \text{m} = \boxed{102.5\ \text{nm}}$$

$$\lambda_3 = \frac{1}{1.097\ 373\ 2 \times 10^7\ \text{m}^{-1}}\left(\frac{4^2}{4^2 - 1}\right) = 9.720 \times 10^{-8}\ \text{m} = \boxed{97.20\ \text{nm}}$$

     (b)   These wavelengths are all in the $\boxed{\text{far ultraviolet region}}$ of the spectrum.

28.2   (a)   The wavelengths in the Paschen series of hydrogen are given by $1/\lambda = R_H(1/3^2 - 1/n^2)$, where $n = 4, 5, 6, \ldots$, and the Rydberg constant is $R_H = 1.097\ 373\ 2 \times 10^7$ m^{-1}. This can also be written as $\lambda = (1/R_H)[9n^2/(n^2 - 9)]$ so the first three wavelengths in this series are

$$\lambda_1 = \frac{1}{1.097\ 373\ 2 \times 10^7\ \text{m}^{-1}}\left[\frac{9(4)^2}{4^2 - 9}\right] = 1.875 \times 10^{-6}\ \text{m} = \boxed{1\ 875\ \text{nm}}$$

$$\lambda_2 = \frac{1}{1.097\ 373\ 2 \times 10^7\ \text{m}^{-1}}\left[\frac{9(5)^2}{5^2 - 9}\right] = 1.281 \times 10^{-6}\ \text{m} = \boxed{1\ 281\ \text{nm}}$$

$$\lambda_1 = \frac{1}{1.097\ 373\ 2 \times 10^7\ \text{m}^{-1}}\left[\frac{9(6)^2}{6^2 - 9}\right] = 1.094 \times 10^{-6}\ \text{m} = \boxed{1\ 094\ \text{nm}}$$

     (b)   These wavelengths are all in the $\boxed{\text{infrared region}}$ of the spectrum.

28.13   The energy absorbed by the atom is

$$E_{\text{photon}} = E_f - E_i = 13.6\ \text{eV}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

     (a) $E_{\text{photon}} = 13.6\ \text{eV}\left(\dfrac{1}{2^2} - \dfrac{1}{5^2}\right) = \boxed{2.86\ \text{eV}}$

     (b) $E_{\text{photon}} = 13.6\ \text{eV}\left(\dfrac{1}{4^2} - \dfrac{1}{6^2}\right) = \boxed{0.472\ \text{eV}}$

28.15 To ionize the atom, it is necessary that $n_f \rightarrow \infty$. The required energy is then

$$\Delta E = E_f - E_i = -13.6 \text{ eV} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) = -13.6 \text{ eV} \left(\frac{1}{\infty} - \frac{1}{n_i^2} \right) = \frac{13.6 \text{ eV}}{n_i^2}$$

(a) If $n_i = 1$, the required energy is $\Delta E = \dfrac{13.6 \text{ eV}}{1^2} = \boxed{13.6 \text{ eV}}$

(b) If $n_i = 3$, $\Delta E = \dfrac{13.6 \text{ eV}}{3^2} = \boxed{1.51 \text{ eV}}$

28.27 In the $3d$ subshell, $n = 3$ and $\ell = 2$. The 10 possible quantum states are

$n = 3$	$\ell = 2$	$m_\ell = +2$	$m_s = +\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = +2$	$m_s = -\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = +1$	$m_s = +\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = +1$	$m_s = -\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = +0$	$m_s = +\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = +0$	$m_s = -\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = -1$	$m_s = +\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = -1$	$m_s = -\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = -2$	$m_s = +\frac{1}{2}$
$n = 3$	$\ell = 2$	$m_\ell = -2$	$m_s = -\frac{1}{2}$

28.28 (a) For a given value of the principal quantum number n, the orbital quantum number ℓ varies from 0 to $n-1$ in integer steps. Thus, if $n = 4$, there are $\boxed{4}$ possible values of ℓ: $\ell = 0, 1, 2,$ and 3.

(b) For each possible value of the orbital quantum number ℓ, the orbital magnetic quantum number m_ℓ ranges from $-\ell$ to $+\ell$ in integer steps. When the principal quantum number is $n = 4$ and the largest allowed value of the orbital quantum number is $\ell = 3$, there are $\boxed{7}$ distinct possible values for m_ℓ. These values are: $m_\ell = -3, -2, -1, 0, +1, +2,$ and $+3$.

28.30   (a)   The electronic configuration for nitrogen $(Z = 7)$ is $\boxed{1s^2 2s^2 2p^3}$.

(b)   The quantum numbers for the 7 electrons can be:

1s states	$n = 1$	$\ell = 0$	$m_\ell = 0$	$m_s = +\frac{1}{2}$
				$m_s = -\frac{1}{2}$
2s states	$n = 2$	$\ell = 0$	$m_\ell = 0$	$m_s = +\frac{1}{2}$
				$m_s = -\frac{1}{2}$
2p states	$n = 2$	$\ell = 1$	$m_\ell = -1$	$m_s = +\frac{1}{2}$
				$m_s = -\frac{1}{2}$
			$m_\ell = 0$	$m_s = +\frac{1}{2}$
				$m_s = -\frac{1}{2}$
			$m_\ell = 1$	$m_s = +\frac{1}{2}$
				$m_s = -\frac{1}{2}$

28.32   (a)   For $n = 1$, $\ell = 0$ and there are $2(2\ell + 1)$ states $= 2(1) = \boxed{2}$ sets of quantum numbers.

(b)   For $n = 2$, $\ell = 0$ for $2(2\ell + 1)$ states $= 2(0+1) = 2$ sets
and $\ell = 1$ for $2(2\ell + 1)$ states $= 2(2+1) = 6$ sets
total number of sets $= \boxed{8}$

(c)   For $n = 3$, $\ell = 0$ for $2(2\ell + 1)$ states $= 2(0+1) = 2$ sets
and $\ell = 1$ for $2(2\ell + 1)$ states $= 2(2+1) = 6$ sets
and $\ell = 2$ for $2(2\ell + 1)$ states $= 2(4+1) = 10$ sets
total number of sets $= \boxed{18}$

(d)   For $n = 4$, $\ell = 0$ for $2(2\ell + 1)$ states $= 2(0+1) = 2$ sets
and $\ell = 1$ for $2(2\ell + 1)$ states $= 2(2+1) = 6$ sets
and $\ell = 2$ for $2(2\ell + 1)$ states $= 2(4+1) = 10$ sets
and $\ell = 3$ for $2(2\ell + 1)$ states $= 2(6+1) = 14$ sets
total number of sets $= \boxed{32}$

(e)   For $n = 5$, $\ell = 0$ for $2(2\ell + 1)$ states $= 2(0+1) = 2$ sets
and $\ell = 1$ for $2(2\ell + 1)$ states $= 2(2+1) = 6$ sets
and $\ell = 2$ for $2(2\ell + 1)$ states $= 2(4+1) = 10$ sets
and $\ell = 3$ for $2(2\ell + 1)$ states $= 2(6+1) = 14$ sets
and $\ell = 4$ for $2(2\ell + 1)$ states $= 2(8+1) = 18$ sets
total number of sets $= \boxed{50}$

For $n = 1$: $2n^2 = 2$ For $n = 2$: $2n^2 = 8$
For $n = 3$: $2n^2 = 18$ For $n = 4$: $2n^2 = 32$
For $n = 5$: $2n^2 = 50$

Thus, the total number of sets of quantum states agrees with the $2n^2$ rule.