General Chemistry I  Worksheet #7a and 7  Quantum Mechanics, Quantum Numbers, Orbitals, electron configurations

1. Define quantum:

A discrete amount of energy possible (whole numbers). A set amount/allowable steps.

2. a. Police often monitor traffic with “K-band” radar guns that operate at 22.235 GHz. Find the wavelength of this radiation in nm. In what part of the electromagnetic spectrum is this radiation found?

\[ c = 2.998 \times 10^8 \text{ m/s} \]
\[ v = (22.235 \text{ GHz})(1\times10^9 \text{ Hz}) = 2.2235 \times 10^{10} \text{ Hz} \]
\[ \lambda = \frac{c}{v} = \frac{2.998 \times 10^8 \text{ m/s}}{2.2235 \times 10^{10} \text{ Hz}} = 0.013483 \text{ m} \]

(1 GHz) \[ \lambda = 0.013483 \text{ m}(1 \text{ nm}) = 1.3483 \times 10^7 \text{ m} \]

(1x10^{-9} \text{ m})

(this is the radiofrequency range)

b. Calculate the energy of this radiation.

\[ E = h\nu = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{0.013483 \text{ m}} = 1.4733 \times 10^{-23} \text{ J} \]

3. a. Sketch the Bohr model of the atom. Label the subatomic particles.

The center of the atom (nucleus) contains the protons and neutrons. The electrons orbit around the nucleus in defined orbits. \( n = 1 \) is the ground state (lowest energy).
b. Using this model, explain why hydrogen produces a line spectrum and not a continuous spectrum.

The electron can only go to certain distances away from the nucleus; it can only be in certain orbitals. When the electron is excited (energy is put in either via electricity or fire, or whatever), it jumps up to a higher orbit (one further from the nucleus). The electron doesn’t stay in this higher energy configuration for long, it soon falls back down to a lower energy orbital (falls back to a orbital closer to the nucleus). When the electrons falls down to a lower energy orbital, it gives off energy (often as visible light; a particular color corresponds to a particular energy). Since the electron can only be in certain orbitals, it can only give off only certain energies (certain colors of light). So the spectrum appears as choppy lines, not a smeared out continuum of colors of the rainbow. Energy is quantized (can only come in certain, defined sized packets).

c. As opposed to the Bohr model of the atom, the Schrödinger model does not have to impose quantization, it occurs naturally as a result. Why?

Schrödinger modeled the electron in the atom as a particle and wave. Using wave mechanics (math to describe waves) gives results that naturally lead to quantization because standing waves can only exist certain distances away from the nucleus because the standing wave can only have whole number multiples of half-wavelengths (or they will cancel themselves out!).

4. What is the maximum number of electrons in an atom that can have these quantum numbers?

(Electrons typically pair up when possible… usually; opposite spins generate opposite magnetic fields. Like a ‘north’ end and a ‘south’ end of a magnet-pairing up- this minimizes the repulsion between two negatively charged species)

a. \( n = 2, \ l = 1 \)

\( p \)-orbital

\( m_\ell = 1, 0, -1 \)

\( p \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \)

six electrons

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

\( d \)-orbital

\( m_\ell = 2, 1, 0, -1, -2 \)

\( d \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \)

ten electrons

c. \( n = 5, m_\ell = +1 \)

\( p, d, f, g \)-orbitals

each can hold 2 electrons in \( m_\ell = +1 \)

(doesn’t apply to s-orbitals;\( m_\ell = 0 \))

eight electrons
5. What are the four quantum numbers for a spin up valence electron with Z = 19?

$$Z = 19 \Rightarrow K \text{ (potassium)}$$

K = 4s¹

$$n = 4, l = 0, m_l = 0, m_s = +1/2$$

6. Can an orbital with the following quantum numbers exist? If not, why not? OR If so, what orbital, suborbital and spin is it?

a. \(n = 3, l = 1, m_l = 0, m_s = -1/2\) (spin down)

\[
\begin{array}{ccc}
& & \\
\downarrow & \downarrow & \downarrow \\
3 & p & -1 & 0 & +1
\end{array}
\]

b. \(n = 2, l = 2, m_l = 0, m_s = -1/2\)

\[
\begin{array}{ccc}
& & \\
\downarrow & \downarrow & \\
2 & d & \\
\end{array}
\]

The value of \(l\) must be \(n-1\) at most, this orbital cannot exist (there is no 2d orbital).

7. Draw the shapes (including the orientation) of the \(s, p\) and \(d\)-orbitals and give the complete name (x, y, z notation).

8. Using the Aufbau principle, write the expected electron configurations and the orbital (‘box’) diagrams following the Pauli exclusion principle and Hund’s rule for each of the following atoms (watch out for any anomalies):

a. S \(1s^2 2s^2 2p^6 3s^2 3p\uparrow \downarrow \)

b. Ca \(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2\)

c. Cr \(1s^2 2s^2 2p^6 3s^2 3p^6 4s\uparrow 3d \uparrow \downarrow \downarrow \downarrow \)