Name: ___________________

INTRODUCTION TO QUANTUM MECHANICS

1. Define the terms:

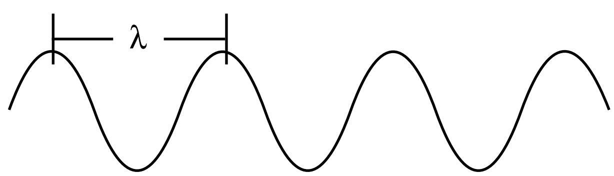
electromagnetic radiation

Electromagnetic radiation is the term used to refer to all types of radiant energy. Electromagnetic radiation is characterized by wave properties. Visible light is one example of a type of electromagnetic radiation. Other examples include, infrared radiation, ultraviolet light, X-rays and microwaves.
Electromagnetic radiation has four characteristic wavelike properties. They are; wavelength, frequency, amplitude and velocity.

wave

A continuously repeating, periodic pulsing of matter or of electric and magnetic fields.

wavelength
Wavelength is defined as the distance between any two repeating points on a wave. In the wave below, the wavelength, λ, measures the length from peak to peak of the wave. The units for wavelength are units of length.

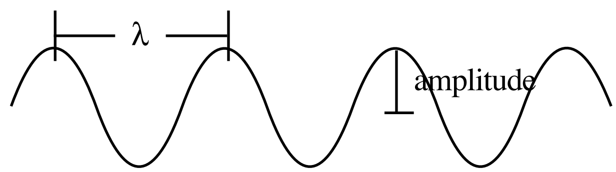


frequency

Frequency is the number of times in a second a single complete wavelength passes a particular point in space. The units of frequency are
$\frac{\text{waves (cycles)}}{\text{s}}$ **or Hz (hertz). The symbol for frequency is υ.**

amplitude

Amplitude is the distance from the highest, or lowest, point on a wave and the midpoint (center) of the wave. The units of amplitude are units of length.



velocity

The velocity of a wave is determined from the product of its wavelength and its frequency.

$$\lambda \cdot \upsilon = \text{velocity}$$

The velocity of electromagnetic radiation in a vacuum is approximately 3.00×10^{8} m/s is given the symbol *c*.

2. Write the equation that describes the mathematical relationship between wavelength and frequency using the standard symbol for each quantity.

$$\lambda \cdot \upsilon = c$$

λ = is the wavelength of the electromagnetic radiation
υ = is the frequency of the electromagnetic radiation
c = is the speed of light

3. Calculate the frequency of light which has a wavelength of 6.7×10^{-5} cm.

$$\lambda \cdot \upsilon = c$$

$$\upsilon = \frac{c}{\lambda}$$

$$\upsilon = \left(\frac{3.00 \times 10^{8}\ \text{m/s}}{6.7 \times 10^{-5}\ \text{cm}}\right)\left(\frac{100\ \text{cm}}{1\ \text{m}}\right)$$

$$= 4.5 \times 10^{14}\ \text{s}^{-1}$$

4. Define quantization. What is a quantum of matter? What is a quantum of light (radiant energy)?

Quantization is the idea that matter or light is composed of discrete packets or units. For example, matter is composed of discrete units called atoms. We do not find fractions of atoms. The quantum of light is called a photon. It is an energy packet of light.

5. Write the relationship between the energy of a photon of light and its frequency.

$$\mathbf{E = h\upsilon}$$

According to Planck's experiment, on the emission of light from a heat object, the energy of the light was directly proportional to the frequency of the light.

6. Calculate the energy of a photon of orange light with a frequency of 5.0×10^{14} sec^{-1}.

$\mathbf{E = h\upsilon}$

$\mathbf{(6.626 \times 10^{-34}\ J{\cdot}s)\ (5.0 \times 10^{14}\ s^{-1}) = 3.3 \times 10^{-19}\ J}$

7. Calculate the energy of a mol of photons of orange light with a frequency of 5.0×10^{14} sec^{-1}.

From above we know one photon has 3.3×10^{-19} J.

$$\mathbf{(3.3 \times 10^{-19}\ \frac{J}{photon})\ (6.02 \times 10^{23}\ \frac{photon}{mole}) = 2.0 \times 10^{5}\ \frac{J}{mole}}$$

8. Calculate the energy of a photon of light with a wavelength of 425 nm.

$$\mathbf{\lambda \cdot \upsilon = c \quad ; \quad \upsilon = \frac{c}{\lambda} \quad ; \quad E = h\upsilon}$$

$$\mathbf{E = h\left(\frac{c}{\lambda}\right) = (6.626 \times 10^{-34}\ J{\cdot}s)\ \frac{3.00 \times 10^{8}\ \frac{m}{s}}{425 \times 10^{-9}\ m} = 4.68 \times 10^{-19}\ J}$$

9. The energy required to break the oxygen-oxygen bond in O_2 is 496 $\frac{\text{kJ}}{\text{mol}}$. Calculate the minimum wavelength of light that can break the oxygen-oxygen bond.

$$\frac{\mathbf{496 \times 10^{3}\ J}}{\mathbf{6.023 \times 10^{23}\ molecules}} = \mathbf{8.24 \times 10^{-19}\ \frac{J}{molecule}}$$

$$\mathbf{E = \frac{hc}{\lambda}\ ;\ \lambda = \frac{hc}{E}}$$

$$\lambda = \frac{\mathbf{6.626 \times 10^{-34}\ J{\cdot}s \cdot 2.998 \times 10^{8}\ \frac{m}{s}}}{\mathbf{8.24 \times 10^{-19}\ \frac{J}{molecule}}}$$

$$\mathbf{= 2.41 \times 10^{-7}\ m = 241\ nm}$$

10. List the postulates about the behavior in atoms that Bohr invoked to explain the occurrence of emission lines and their associated energies in the hydrogen spectrum.

1. **The electron in a hydrogen atom moves around the nucleus in a circular orbit with a particular radius.**
2. **An electron can only exist at certain integral distances from the nucleus. The further the electron is from the nucleus the greater the energy of the electron. Since only orbits of certain radii are allowed the electron can have only certain values of energy.**
3. **If no light is incident on an atom the electron remains indefinitely in a particular orbit. The electron is said to exist in a particular energy state.**
4. **When light is incident on the atom the electron can absorb energy and is moved to an orbit further from the nucleus. The electron is said to be excited to a higher energy level. The frequency of the light which excites the electron from one energy level to another is exactly equal to the difference in energy of the two levels.**
5. **When an electron falls from a high energy level, an orbit far away from the nucleus, to a lower energy level, an orbit close to the nucleus, light is emitted. The energy of the light emitted is equal to the difference in energy between the two energy levels.**

11. Briefly describe the difference between the continuous model of matter and quantized view of matter.

In the continuous model, matter (or energy) can be subdivided to any degree desired. For example, any mass of a substance could be obtained if one had a balance sensitive enough to measure it.
In the quantized model, however, matter (or energy) comes in discreet, indivisible particles. Only whole number multiples of these quantized particles can be obtained. For example, one could have one atom of carbon, or one mole of atoms, but one could never have 0.5 atoms.

12. Which color of light is higher energy, blue of red? Explain.

Blue light is higher in energy than red. Blue light has a shorter wavelength and higher frequency than red light. Energy of light is calculated using the equation: $E = h\upsilon$. Higher frequency results in higher energy.

13. An argon laser emits light with a wavelength of 488 nm. Calculate the energy of a photon of light emitted from the laser.

$$E = h\upsilon \;;\; \upsilon = \frac{c}{\lambda}$$

$$E = (6.626 \times 10^{-34}\ J\ s)\left(\frac{2.9979 \times 10^{8}\ m\ s^{-1}}{488 \times 10^{-9}\ m}\right) = 4.07 \times 10^{-19}\ J$$

14. Calculate the energy and the wavelength of light emitted when an electron in an excited hydrogen atom falls from the n = 5 level to the n = 1 level.

$$E = -R_H\left(\frac{1}{n^2}\right) \;;\; E_f - E_i = -R_H\left(\frac{1}{n_f^2}\right) - (-R_H\left(\frac{1}{n_i^{\ 2}}\right)) \;;$$

$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) = 2.18 \times 10^{-18}\ J\left(\frac{1}{5^2} - \frac{1}{1^2}\right) = -2.09 \times 10^{-18}\ J$$

$E = h\upsilon \;;\; \upsilon = \frac{c}{\lambda}$ **therefore, E = h Error!) ; λ = h Error!)**

$$\lambda = (6.626 \times 10^{-34}\ J\ s)\left(\frac{2.9979 \times 10^{8}\ m\ s^{-1}}{2.09 \times 10^{-18}\ J}\right) = 9.51 \times 10^{-8}\ m \;;\; 95.1\ nm$$

The wavelength of the photon is 0.951 nm and its energy is -2.09×10^{-18} J.

15. An electron initially in the n = 2 energy level in a hydrogen atom absorbs a photon of light with a frequency of $6.167 \times 10^{14}\ s^{-1}$. Calculate the new energy level the electron will occupy.

$E = h\upsilon$

$E = (6.626 \times 10^{-34}\ J\ s)(6.167 \times 10^{14}\ s^{-1}) = 4.086 \times 10^{-19}\ J$

$$E = \Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$-\frac{1}{n_f^2} = \frac{E}{R_H} - \frac{1}{n_i^2}$$

$$-\frac{1}{n_f^2} = \frac{4.086 \times 10^{-19}\ J}{2.18 \times 10^{-18}\ J} - \frac{1}{2^2} = -0.06257$$

$n_f^2 = 15.98$

$n_f = 4$; The electron is in the n = 4 energy level.

16. Will a photon of light of wavelength 480 nm excite an electron in the hydrogen atom from the n = 1 level to the n = 2 level? Explain.

$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= 2.18 \times 10^{-18}\ J\left(\frac{1}{1^2} - \frac{1}{2^2}\right) = 1.635 \times 10^{-18}\ J$$

$$E = h\upsilon\ ;\ \upsilon = \frac{c}{\lambda}$$

$$E = h\frac{c}{\lambda}\ ;\ \lambda = h\frac{c}{E}$$

$$\lambda = (6.626 \times 10^{-34}\ J\ s)\left(\frac{2.9979 \times 10^{8}\ m\ s^{-1}}{1.635 \times 10^{-18}\ J}\right) = 1.21 \times 10^{-7}\ m$$

$$= 121\ nm$$

A photon of wavelength 480 nm cannot excite a hydrogen atom from the n = 1 to the n = 2 level. This transition requires *exactly* 1.635×10^{-18} J of energy. This corresponds to a wavelength of 121 nm. A photon of wavelength 480 nm does not nave enough energy to excite an electron in a hydrogen atom from the ground state to the *n* = 2 energy level.

17. Define the terms *excited state*, *ground state* and *ionization energy*.

excited state: When an electron occupies an energy level other than the n = 1 (ground state.)

ground state: When an electron is in the n = 1 energy level.

ionization energy: The energy required to remove an electron from an atom when the electron is in its ground state.

18. Calculate the ionization energy for the hydrogen atom using the Bohr model.

$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\Delta E = 2.18 \times 10^{-18}\ J\left(\frac{1}{1^2} - \frac{1}{\bullet}\right) \qquad \text{Note: } \frac{1}{\bullet} \rightarrow 0$$

Ionization energy = 2.18×10^{-18} J

20.What was the importance of De Broglie's postulate of the wave nature of matter?

De Broglie postulated that if light could exhibit both wave and particle characteristics then why could not a particle, such as an electron, exhibit both particle and wave character? De Broglie suggested the wavelength of a particle could be calculated using the equation;

$$\lambda = \frac{h}{m\upsilon}$$

21.Define the uncertainty principle and explain its importance. **See Appendix III for recommended demonstration, video, or computer resources**

The Heisenberg uncertainty principle suggested there is a limit to how precisely it is possible to know both the position and the energy (momentum) of a particle. This idea was used to expand the picture of an electron in an atom. The uncertainty principle suggests it is incorrect to think of the electron as moving in a well-defined orbit (path) around the nucleus. It is better to think of the electron as a diffuse cloud of 'electron density'. The density and shape of the cloud reflect the probability of finding the electron at a particular location in the region around the nucleus.

22.What are the critical shortcomings of the Bohr model of the atom? How does the quantum mechanical description of the atom overcome these shortcomings.

The Bohr model could not successfully explain the atomic spectrum of an atom with more than one electron. The electron in the Bohr model was assumed to behave as a particle and only one quantum number was required to define the energy and the location of the electron. In the quantum mechanical model of the atom the electron is described as a wave. In this model three quantum numbers are required to specify the electron.

23.Define the term *orbital*.

Orbital is the term used to describe the region in space where an electron is most likely to be found.

24.Define the three quantum numbers, principal, azimuthal and magnetic using mathematical relationships. How are these quantum numbers related to the terms shell, subshell and orbital?

***n* is the principal quantum number. It identifies the shell the electron occupies and the energy of the electron. *n* can have only whole number values from 1 to infinity.**
The azimuthal quantum number, *l*, describes the shape of the orbital and it identifies the subshell occupied by the electron. The values of *l* are dependent on the *n* shell. *l* values are whole numbers ranging from zero to *n*-1.

The magnetic quantum number, *m*, describes the orientation of the orbital. The combination of these three quantum numbers uniquely identifies the orbital an electron occupies. Values for *m* are determines by the l (subshell) value. *m* values must be integers and range from -*l* to +*l*, including zero.

25. How are these quantum numbers related to energy, shape and orientation of the electron?

***n* determines the energy of the electron.**
***l* determines the shape of the orbital occupied by the electron.**
***m* determines the orientation of the occupied orbital in space.**

26. Prepare a table of the possible sets of quantum numbers for an electron in the $n = 1, 2$, 3 and 4 levels. **See Appendix III for recommended demonstration, video, or computer resources**

Principle quantum # n	Azimuthal quantum # l	Magnetic quantum # m_l
1	**0**	**0**
2	**0**	**0**
	1	**+1, 0, –1**
3	**0**	**0**
	1	**+1, 0, –1**
	2	**+2,+1, 0, –1,–2**
4	**0**	**0**
	1	**+1, 0, –1**
	2	**+2,+1, 0, –1,–2**
	3	**+3,+2,+1, 0, –1,–2,–3**

27. Which of the following combinations of quantum numbers is allowed for an electron in a hydrogen atom?

i) $n = 2; l = 2; m_l = 0$ **not allowed**

ii) $n = 3; l = 1; m_l = -1$ **allowed**

iii) $n = 5; l = 0; m_l = +1$ **not allowed**

iv) $n = 1; l = 0; m_l = 0$ **allowed**

28. How many orbitals are available with the following combination of quantum numbers?

i) $n = 2; l = 0$ **1 orbital (2s)**

ii) $n = 3; l = 2$ **5 orbitals (the 3d's)**

iii) $n = 4; l = 3$ **7 orbitals (the 4f's)**

29. What is the subshell designation for an electron with the following set quantum numbers?

(i) $n = 1; l = 0$ ***1s***

(ii) $n = 3; l = 2; m_i = -1$ ***3d***

(iii) $n = 4; l = 3$ ***4f***

(iv) $n = 2; l = 0; m_i = 0$ ***2s***

30. What is the maximum number of electrons in a

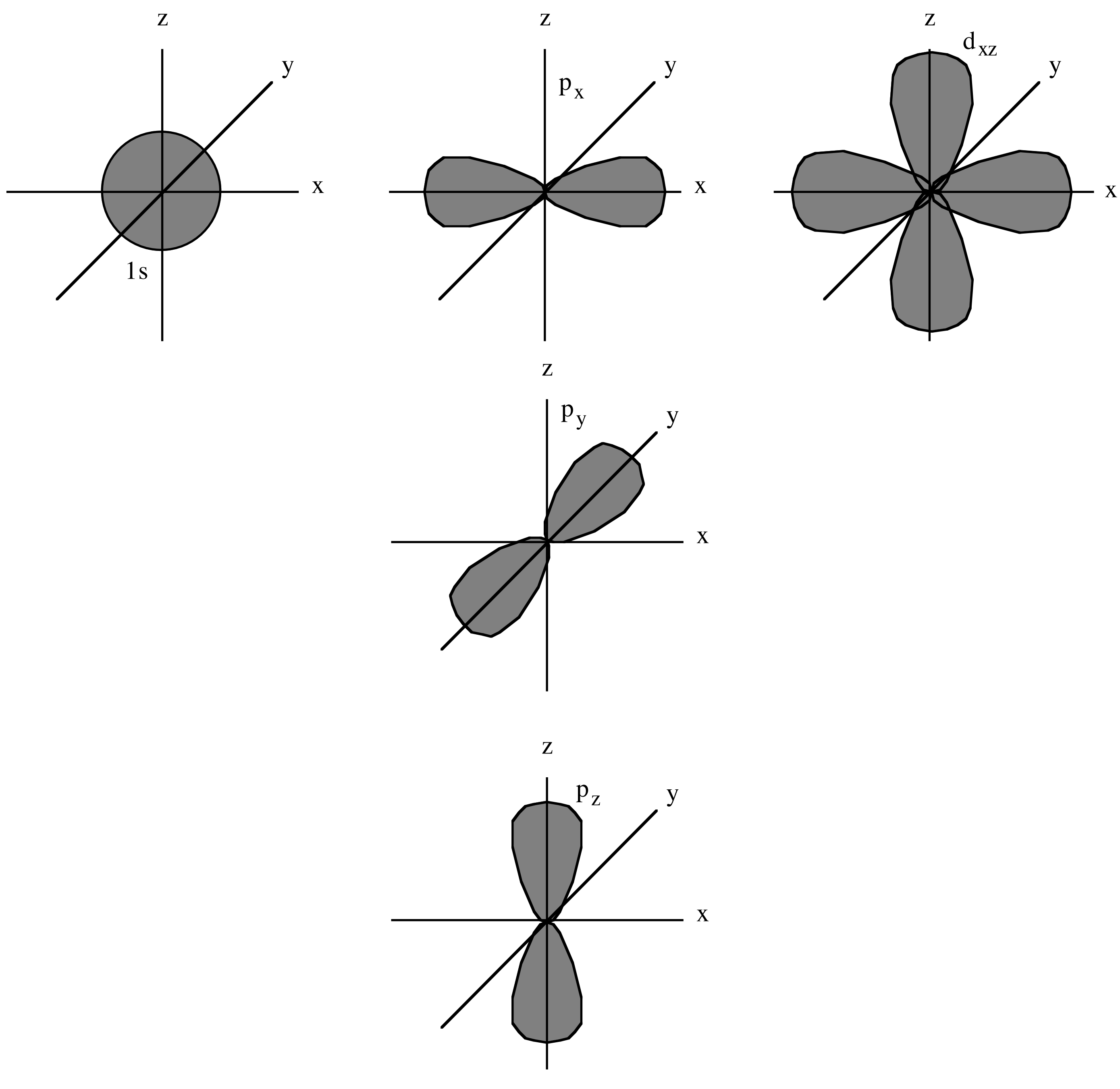
i) *2s* subshell **2**

ii) *3d* subshell **10**

iii) $n = 4$ shell **32**

iv) *2p* orbital **2**

31. Draw the contour representation that depicts the shape of an *s* orbital, a p_x, p_y, p_z, orbital and a d_{xz} orbital. **See Appendix III for recommended demonstration, video, or computer resources.**



32. How does the contour representation of a *1s* orbital differ from that of a *2s* orbital?

The *2s* orbital is larger than the *1s* orbital. The *2s* orbital also contains a node which is not present in the *1s* orbital.