AP Chemistry Review Sheet -- Chapter 6 -- Atomic Structure

Target #1: I can use $\lambda\nu = c$ to solve for the wavelength or the frequency of a wave.

1. Circle the possible units of frequency and underline the units of wavelength. Hint: Some of the units may not be for either wavlength or frequency. Do not circle or underline these!

   <u>cm</u> cm/sec s^{-1} Hertz <u>m</u> 1/sec g/cm^3

2. Circle all of the following statements which are TRUE statements.
   a) Wavelength is directly proportional to frequency. ⟶ inversely related
   b) As the wavelength increases, the frequency also ~~increases~~. ⟶ ν decreases
   c) The shorter the wavelength of a wave, the higher its frequency.
   d) A light wave will travel faster if its wavelength is shorter. ⟶ same speed (c)
   e) The speed of light is 3.0×10^8 m/s.
   f) The speed of sound is faster than the speed of light. sound travels a lot slower

3. Assume that a wave has a wavlength of 25.0 cm. Calculate its frequency in Hertz.

   25.0 cm = 0.250 m

   $$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{0.250 \text{ m}} = 1.20 \times 10^9 \text{ Hz}$$

4. A radiowave has a frequency of 5.4×10^{10} s^{-1}. Calculate the wavelength of the radiowave (in meters).

   $$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8 \text{ m/s}}{5.4 \times 10^{10} \text{ s}^{-1}} = 5.6 \times 10^{-3} \text{ meters}$$

5. How many minutes would it take a radio wave to travel from the planet Venus to Earth? (The average distance from Venus to Earth is 28 million miles.)

   $$X \text{ min} = \frac{28{,}000{,}000 \text{ miles}}{} \left| \frac{1.609 \text{ km}}{1 \text{ mile}} \right| \frac{10^3 \text{ m}}{1 \text{ km}} \left| \frac{1 \text{ sec}}{3.00 \times 10^8 \text{ m}} \right| \frac{1 \text{ min}}{60 \text{ sec}} = 2.5 \text{ min}$$

Target #2: I can describe the component parts of the electromagnetic spectrum.

6. List the various parts of the electromagnetic spectrum in order from highest energy to lowest energy.

GAMMA RAYS	X-RAYS	UV	VISIBLE (VIBGYOR)	IR	MICRO-WAVES	FM RADIO	AM RADIO

7. List the order of visible light from longest wavelength to shortest wavelength. Also, what is the wavelength range of visible light?

   ROYGBIV

   400 nm ⟶ 700 nm
   V R

8. Circle the letters of all TRUE statements below:
   a) Blue light has more energy than orange light.
   b) Microwaves have a higher frequency than UV light.
   c) Gamma rays move at a faster speed than X-rays. SAME
   d) Yellow light is more energetic than FM radio waves.

Target #3: I can explain Max Planck's quantum theory and its significance to understanding the structure of matter.

The fact that different colors are produced lends support to QUANTUM THEORY.

9. Max Planck assumed that energy can be either released or absorbed in by atoms only in discrete "chunks" of some minimum size. What name did he give to these chunks?

PHOTONS or QUANTA

10. What equation did Planck use to help explain his idea of quantized energy? What is the name and value of the constant used in this equation?

$\boxed{E = h\nu}$ Planck's constant ; $h = 6.63 \times 10^{-34}$ J·s

11. Which of the following each pairs is considered "quantized"? Circle your answer.
   a) piano or a violin
   b) a ramp or stairs

12. The temperature of stars is gauged by their colors. For example, red stars have a lower temperature than blue stars. How is this temperature scale consistent with Planck's assumptions regarding quantized energy?

Blue stars release more energy (are hotter) than red stars. The energy of a photon from blue stars in greater, has a larger frequency and thus a shorter wavelength. These discrete amounts of energy released produce diff. colors.

13. Assume that a laser emits light with a frequency of 5.09×10^{14} s^{-1}.
   a) Calculate the energy (in joules) of a photon of this laser light.

$E = h\nu = (6.63 \times 10^{-34} \text{ J·s})(5.09 \times 10^{14} \text{ s}^{-1}) = \boxed{3.37 \times 10^{-19} \text{ J}}$ per photon

   b) If the laser emits 1.8×10^{-4} J of energy during a pulse, how many photons are emitted during the pulse?

$x \text{ photons} = \frac{1.8 \times 10^{-4} \text{ J}}{} \left| \frac{1 \text{ photon}}{3.37 \times 10^{-19} \text{ J}} \right| = \boxed{5.3 \times 10^{14} \text{ photons}}$

Target #4: I can explain how Einstein accounted for the photoelectric effect and its significance to understanding the structure of matter.

14. What is the photoelectric effect? The emission of e- from a metal surface induced by light.

15. Calculate the energy of a photon which has a wavelength of 610 nm.

↳ 6.1×10^{-7} meters

$E = h\nu = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J·s})(3.00 \times 10^{8} \text{ m/s})}{6.1 \times 10^{-7} \text{ m}}$

$E = 3.3 \times 10^{-19}$ J

Target #5: I can explain the origin and difference between a line spectrum and a continuous spectrum. I will also be able to describe Johann Balmer's contribution to calculating wavelengths of the lines in a line spectrum.

16. What is a line spectrum?

A spectrum that contains radiation at only specific wavelengths.

17. What causes the lines on a line spectrum?

The lines are caused by excited e- falling back down to a lower state. As they (the e-) fall, they release energy in the form of light. The various amounts of energy released causes the various lines.

18. Why are the lines on a line spectrum different colors?

Different colors are produced because the emitted photons from "falling" e- are of different energies (due to different energy levels in the atom). Different energy → different ν → different λ → diff. colors

19. How is a continuous spectrum different than a line spectrum?

A continuous spectrum contains a "RAINBOW" of colors due to a wide range of radiation over ALL wavelengths.

20. What was Balmer's contributions to calculating the wavelengths of the lines in a line spectrum?

Balmer showed (mathematically) that the wavelengths of the 4 lines of hydrogen in the line spectrum of hydrogen were related to integers.

$$\boxed{\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)}$$

NO NEED TO MEMORIZE!

Target #6: I can explain the contributions of Neils Bohr to quantum theory. I will also be able to calculate the energies corresponding to each allowed orbit for an electron in the hydrogen atom. I will also be able to calculate the energy difference between two allowed states.

21. Describe Neils Bohr's significant contributions to quantum theory.

(1) In hydrogen atom, only orbits of specific energies are allowed.

(2) Electrons can only be in these "allowed" energy states.

(3) Energy is emitted or absorbed as result of e- changing states.

22. Calculate the energy of a n = 3 orbit in a Bohr hydrogen atom.

$$E_3 = \frac{-2.18\times10^{-18}\,J}{3^2} = \boxed{-2.42\times10^{-19}\,J}$$

23. Calculate the energy of a n = 5 orbit in a Bohr hydrogen atom.

$$E_5 = \frac{-2.18\times10^{-18}\,J}{5^2} = \boxed{-8.72\times10^{-20}\,J}$$

24. Calculate the energy, frequency, and wavlength of a photon released as an electron falls from the n = 5 state to n = 3 state on a Bohr hydrogen atom.

$$\Delta E = E_f - E_i = E_3 - E_5 = -2.42\times10^{-19} - (-8.72\times10^{-20}\,J) = \boxed{-1.55\times10^{-19}\,J}$$

$$\Delta E = h\nu\,;\ \nu = \frac{\Delta E}{h} = \frac{-1.55\times10^{-19}\,J}{6.63\times10^{-34}\,J\cdot s} = \boxed{2.33\times10^{14}\,s^{-1}}$$

$$\lambda = \frac{c}{\nu} = \frac{3.00\times10^{8}\,m/s}{2.33\times10^{14}\,s^{-1}} = \boxed{1.29\times10^{-6}\,m}$$

25. The first line of the Bālmer series occurs at a wavelength of 656.3 nm. What are the values for n_i and n_f for the electrons which cause this line of the Balmer series?

BALMER SERIES automatically implies $n_f = 2$. The first line is a transition of $n=3 \rightarrow n=2$. You could solve mathematically...

$$\Delta E = h\nu = \frac{hc}{\lambda} = \frac{(6.63\times10^{-34}\,J\cdot s)(3.00\times10^{8}\,m/s)}{656.3\times10^{-9}\,m} = -2.18\times10^{-18}\,J\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

SOLVE THIS EQ. FOR n_i ☺

$(n=2)$

26. A photon of UV radiation has an energy of -2.043 X 10^{-18} J. Determine its n_i and n_f.

$n_f = 1$

$$\Delta E = -2.18\times10^{-18}\,J\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$n_i = 1$
$n_f = 4$

$$\Delta E = -2.18\times10^{-18}\,J\left(\frac{1}{1^2} - \frac{1}{n_i^2}\right) = \boxed{-2.09\times10^{-18}\,J}$$

Solve for n_i!! ☺

Target #7: I can explain the work of Louis deBroglie (matter waves) and the contributions that he made to quantum theory. I will be able to calculate the wavelength of a particle given its mass and velocity.

27. Define MATTER WAVE: the term used to describe the wave characteristics of a moving particle

28. Protons can be accelerated to speeds near that of light in particle accelerators. Estimate the wavelength (in nm) of such a proton moving at 2.90 X 10^8 m/s. (Mass of a proton is 1.673 X 10^{-27} kg)

$$\lambda = \frac{h}{m \cdot v}$$

$$\lambda = \frac{6.63 \times 10^{-34}\ \text{J} \cdot \text{s}}{(1.673 \times 10^{-27}\ \text{kg})(2.90 \times 10^{8}\ \text{m/s})} = 1.37 \times 10^{-15}\ \text{meters} \Rightarrow \boxed{1.37 \times 10^{-6}\ \text{nm}}$$

Target #8: I can explain the Heisenberg uncertainty principle.

29. State the Heisenberg Uncertainty Principle.

It is impossible to know simultaneously both the exact momentum & location of an e⁻ in space.

30. What is the significance of the Heisenberg Uncertainty Principle? In other words, how did this principle help us develop our current understanding of the structure of an atom?

Any thoughts of determining the location of an e- were abandoned. The WAVE nature of the e- was recognized. Because of the work of Heisenberg & DeBroglie, we described very precisely the ENERGY of the e- but only spoke of the location of the e- in terms of probabilities.

Target #9: I can explain the concepts of energy shell, subshell, and orbitals as descibed by the quantum-mechanical model of the atom.

31. What is the difference between a $2p_x$ and $2p_y$ orbital?

SAME SHAPE
SAME ENERGY

→ The only difference is their orientation in space. $2p_x$ has e- density along x-axis & $2p_y$ has e- density along y-axis.

N: $1s^2 2s^2 2p^3$

Se: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

Si: $1s^2 2s^2 2p^6 3s^2 3p^2$

32. a) How many p electrons are in an atom of nitrogen? 3
b) How many s electrons are in an atom of silicon? 4
c) How many 3d electrons are in an atom of selenium? 10
d) How many f electrons are in an f subshell which is half-filled? 7
e) How many electrons can each of the following energy levels hold if completely filled?
n=1 2, n=2 8, n=3 18, n=4 32 $2n^2$

33. List three orbitals in an atom of magnesium which are degenerate. $2p_x$, $2p_y$, $2p_z$

34. Consider the following orbitals: $4p_z$, $4p_y$, $4p_x$, $4s$, $4d_{xz}$, $4d_{x2-y2}$
a. Which are not acceptable notations for orbitals?
All acceptable (!)
b. Which are degenerate orbitals in a . . .
i) many electron system? all orbitals in same subshell
ii) single e- system? all orbitals in same energy shell.

35. Which do you think is bigger, the 1s orbital in hydrogen or the 1s orbital in lithium? Explain.
Has less p+ than Li so it won't "pull" on shell as hard.

Target #10: I can write electron configurations and draw orbital diagrams. I can state Hund's Rule and the Pauli Exclusion Principle and I am able to recogize violations of these rules.

36. Write the ground state electron configurations for the following elements:

a) Ge $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

b) Fe $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

c) Zn $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

37. Draw the orbital diagrams for each of the following elements:

a) Al $1s^2 2s^2 2p^6 3s^2 3p^1$; ↑↓ 1s ↑↓ 2s ↑↓ ↑↓ ↑↓ 2p ↑↓ 3s ↑ _ _ 3p

b) Ca ↑↓ 1s ↑↓ 2s ↑↓ ↑↓ ↑↓ 2p ↑↓ 3s ↑↓ ↑↓ ↑↓ 3p ↑↓ 4s

$s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d$ $4s^2$

38. What postulate negated Bohr's idea that e- travel in well defined paths? Please explain.
Heisenberg Uncertainty Principle:

39. Write an electron configuration representing:

a- the ground state of a helium atom. $1s^2$

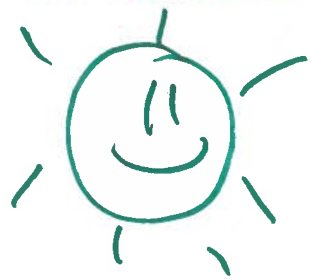
b- an excited state of a helium atom. $1s^1 2s^1$

c- a helium +1 ion. $1s^1$

40. Which of the following orbital diagrams represent excited atoms?

a) ↑ ↑ ↑ ($2p_x$ $2p_y$ $2p_z$) b) ↑↓ (1s) $\boxed{↓}$ (2s) ↑ ↑ ↑ (2p 2p 2p)

e- was excited from 2s up to 2p. ☺



41. Draw an orbital diagram for nitrogen which violates:

a- the Pauli Exclusion Principle. ↑↑ (1s)

All e- have same set

b- Hund's Rule. ↑ ↓ ↑ (5p)

42. Shown below are portions of orbital diagrams representing the ground state electron configurations of certain elements. Which of them violate the Pauli exclusion principle or Hund's Rule?

	Orbital notation	Does it violate the Pauli exclusion principle?	Does it violate Hund's Rule?
A	↑↓ 2p ↑↓ 2p ↑↑ 2p	YES same quantum #'s	NO
B	↑ 2p ↑↓ 2p ↑ 2p	NO	NO
C	↑↑ 2p ↑↓ 2p __ 2p	YES	YES occupy each orbital before pairing up!

Target #11: I can define the four quantum numbers. I will also be able to write a viable set of quantum numbers for a given electron.

43. Write a possible set of 4 quantum numbers for the 5 electrons in a boron atom.

B $1s^2 2s^2 2p^1$

1s: ↑ 1,0,0,+1/2 ; ↓ 1,0,0,-1/2

2s: ↑ 2,0,0,+1/2 ; ↓ 2,0,0,-1/2

2p: ↑ 2,1,1,+1/2

44. Complete the following table:

subshell	n	l	m_l (show all values)
4p	4	1	-1, 0, 1
3d	3	2	-2, -1, 0, 1, 2
3s	3	0	0
5f	5	3	-3, -2, -1, 0, 1, 2, 3

45. Name the type of electron orbital (i.e. s, p, d, or f) designated by each of the following set of quantum numbers.

2p a) 2,1,1 4d b) 4,2,0 1s c) 1,0,0 4p d) 4,1,0

46. Which of the following sets of quantum numbers could NOT occur? Explain.

a) 1, 1, 0, +1/2

b) 2, 1, 0, +1/2

c) 2, 0, 1, -1/2

d) 3, 2, 0, -1/2

47. Arrange the following sets of quantum numbers in order of increasing energy. If they have the same energy, place them together. Assume you have a multi-electron system.

a) 4, 2, -1, +1/2 4d
b) 1, 0, 0, -1/2 1s
c) 3, 1, 1, -1/2 3p
d) 2, 0, 0, +1/2 2s
e) 2, 1, 0, +1/2 2p

1s 2s 2p 3p 4d
B D E C A

48. Assign a set of 4 quantum numbers to each electron in the fluorine atom (9 electrons) by completing the chart below:

$1s^2 2s^2 2p^5$

	n	l	m_l	m_s
$1s^1$	1	0	0	+1/2
$1s^2$	1	0	0	-1/2
$2s^1$	2	0	0	+1/2
$2s^2$	2	0	0	-1/2
$2p^1$	2	1	-1	+1/2
$2p^2$	2	1	0	+1/2
$2p^3$	2	1	1	+1/2
$2p^4$	2	1	-1	-1/2
$2p^5$	2	1	0	-1/2

↑↓ ↑↓ ↑↓ ↑↓ ↑
1s 2s 2p

49. An electron in a hydrogen atom is excited from the ground state to the n = 4 state. Circle "T" for true and "F" for false for each of the statements below. If you think the statement is false, please correct the statement to make it true!

T F a) n = 4 is the first excited state. n = 2 first excited state

T F b) It takes ~~more~~ less energy to ionize (remove the e-) from n = 4 than from ground state.

T F c) The electron is farther from the nucleus (on average) in n = 4 than in ground state.

T F d) The wavelength of light emitted when the electron drops from n = 4 to n = 1 is ~~longer~~ shorter than that from n = 4 to n = 2.

T F e) The wavelength the atom absorbs in going from n = 1 to n = 4 is the same that is emitted as it goes from n = 4 to n = 1.

50. Circle true or false for each of the following.

T F a) In the expression $\Delta E = h\nu$, the $h\nu$ term is called a quantum of energy.

T F b) It is possible for an atom of carbon to have an e- in the $3p_x$ orbital. if excited!

T F c) A 2p orbital in a hydrogen atom is of higher energy than a 2s orbital. degenerate in single e- systems

T F d) The e- configuration for an oxygen atom is the same as the electron configuration as an oxide ion.

T F e) ~~Nine~~ Six electrons in an atom can have both n = 3 and l = 1.

T F f) ~~Six~~ Five electrons in an atom of $\boxed{\text{manganese}}$ can have m_l = 1.

3s —
3p — — —
3d — — — — —

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
2 2 1