

Unit 1, Lesson 02: Answers to Homework

1. Read pages 131 – 133, pages 137 – 138 and pages 142 – 146.
2. Summarize and UNDERSTAND the contributions of Planck, Einstein, de Broglie, Schrodinger and Heisenberg to the quantum mechanical model.
 - a) **Planck**: suggested that energy is quantized- it only comes in certain discrete amounts. The smallest unit of energy is called a “quantum”.
 - b) **Einstein**: suggested that energy has particle-like properties in addition to its wave properties. The smallest unit of energy for electromagnetic radiation (light) is a photon.
 - c) **de Broglie**: all objects have wave-like properties. For large objects, the wave-like properties are negligible, but for tiny particles like electrons, the wave-like motion is significant. Electrons do not move in simple, straight line orbits. They have significant wave motion.
 - d) **Schrödinger (1926)** developed the Quantum Mechanical Model of the atom. It is a mathematical model which accounts for the wave-like motion of electrons and predicts where an electron will be found, 95% of the time. He called these regions “orbitals”.
 - e) **Heisenberg (1927)** realized that we can never know BOTH where an electron is (its position), and where its going (its trajectory). By measuring one of these variables, we change the other..
3. Explain how the quantum mechanical model of the atom differs from Bohr’s model of the atom.
 - a) **Bohr’s planetary model**: all of the electrons in a principal quantum level have exactly the same energy because they are all exactly the same distance from the nucleus. The electrons in a principal quantum level are travelling in a precise, defined, predictable orbit around the nucleus.
 - b) **Quantum Mechanical model**: all of the electrons in a principal quantum level are not the same distance from the nucleus, some are slightly further away. Each principal quantum level (after $n = 1$) can be divided into sub-levels that contain a different type of orbital. The electrons are not found in precise, defined orbits. Rather, we can only describe 3-D regions where electrons are likely to be found, 95% of the time.
4. Explain how an orbital is different from an orbit. Be specific.
 - a) An **orbit** is a precise, defined, predictable path. We can write a mathematical equation to describe the orbit so we know exactly where an object on that orbit (such as an electron) will be, 100 % of the time.
 - b) An **orbital** is a 3-D region where we should look for an electron. The electron will be somewhere within this space 95% of the time. Orbitals are not precise, defined or predictable.
5. How many electrons (maximum) are in quantum level 4? $2n^2 = 32$, When $n=3$ $2n^2 = 18$
6. How many electrons can be designated as 3d? 10, 4s: 2, 5f: 14
7. How many **types** of orbitals are there in quantum level 3? $n = 3$ When $n=4$: $n = 4$
8. How **many orbitals** are there in quantum level 2? $n^2 = 4$ When $n=5$: $n^2 = 25$ When $n=3$: $n^2 = 9$
9. How **many electrons** can be held in quantum level 5? $2n^2 = 50$ When $n=1$ $2n^2 = 2$
10. How **many orbitals** are there in quantum level 1? $n^2 = 1$, When $n=4$: $n^2 = 16$
11. How many **types** of orbitals are there in quantum level 5? $n = 5$ When $n=2$: 2, When $n=6$: 6
12. How many electrons can be designated 2p? 6, 4p: 6, 5s: 2, 6f: 14
13. Circle the orbitals which **do not** exist: ~~2p~~ 5s ~~4f~~ ~~5p~~ 3d ~~3s~~ 4d 4s

14. Questions 6, 7 on page 145.

Standard Format	Orbital Diagram	Condensed Format
${}_{11}\text{Na} \quad 1s^2 2s^2 2p^6 3s^1$		$[\text{Ne}] 3s^1$
${}_{12}\text{Mg} \quad 1s^2 2s^2 2p^6 3s^2$		$[\text{Ne}] 3s^2$
${}_{13}\text{Al} \quad 1s^2 2s^2 2p^6 3s^2 3p^1$		$[\text{Ne}] 3s^2 3p^1$
${}_{14}\text{Si} \quad 1s^2 2s^2 2p^6 3s^2 3p^2$		$[\text{Ne}] 3s^2 3p^2$
${}_{15}\text{P} \quad 1s^2 2s^2 2p^6 3s^2 3p^3$		$[\text{Ne}] 3s^2 3p^3$
${}_{16}\text{S} \quad 1s^2 2s^2 2p^6 3s^2 3p^4$		$[\text{Ne}] 3s^2 3p^4$
${}_{17}\text{Cl} \quad 1s^2 2s^2 2p^6 3s^2 3p^5$		$[\text{Ne}] 3s^2 3p^5$
${}_{18}\text{Ar} \quad 1s^2 2s^2 2p^6 3s^2 3p^6$		$[\text{Ne}] 3s^2 3p^6$

15. Questions 10, 11, 12 and 13 on page 150.

- Question 10
- $[\text{Ne}] 3s^1$ is sodium (3^{rd} period and 1 valence electron, so group I)
 - $[\text{He}] 2s^2$ is beryllium (2^{nd} period and 2 valence electrons, so Group II)
 - $[\text{Kr}] 5s^2 4d^{10} 5p^5$ is iodine (5^{th} period and $(2+10+5) = 17$ valence electrons, so Group 17)

- Question 11
- Group II, Period 4: ${}_{20}\text{Ca} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
 - Noble gas in period 6: ${}_{86}\text{Rn} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$
 - Group 12, Period 4: ${}_{30}\text{Zn} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
 - Group 16, Period 2: ${}_{8}\text{O} \quad 1s^2 2s^2 2p^4$

- Question 12
- $s^2 d^1$ (Group 3 or IIIA) Sc, Y, La and Ac
 - $s^2 p^3$ (Group 15 or VB) N, P (including elements with “d” electrons, add As, Sb, Bi)
 - $s^2 p^6$ (Group 18 or VIII) Ne, Ar (including elements with “d” electrons, add Kr, Xe and Rn)

${}_{19}\text{K}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ = $[\text{Ar}]4s^1$	
${}_{28}\text{Ni}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$ = $[\text{Ar}] 4s^2 3d^8$	
${}_{82}\text{Pb}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$ = $[\text{Xe}] 6s^2 4f^{14} 5d^{10} 6p^2$	

16. Write the predicted and actual (experimentally determined) electron configurations for Mo, Ag and Au.

Predicted Electron Configuration	Actual Electron Configuration
${}_{42}\text{Mo}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^4$	${}_{42}\text{Mo}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^5$
${}_{47}\text{Ag}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^9$	${}_{47}\text{Ag}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^{10}$
${}_{79}\text{Au}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^9$	${}_{79}\text{Au}$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1 4f^{14} 5d^{10}$

17. If valence electrons are found in energy levels further from the nucleus than expected, then electrons are in an excited state. State whether the following electron configurations show electrons in their ground state or an excited state:

- $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2$ excited
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$ ground state
- $1s^2 2s^2 2p^6 3p^2$ excited
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^2$ excited
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 6s^1$ excited
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$ ground state
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^4$ excited