Unit #1, Chapter 3 Outline
Atoms, Electrons, and Periodic Trends

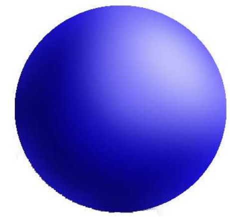
Lesson	Topics Covered	Reference	Homework/Assignments
1	**Review of Atomic Models** • features of models from Dalton, Thomson, Rutherford, Bohr and Chadwick **Review of Atomic Structure** • protons, neutrons and electrons • atomic number (Z) • mass number (A) • isotopes • average atomic mass • ions • standard format	**Note**: Atomic Models from Dalton to Chadwick **Text**: p. 118 – 130	1. Complete all notes and homework titled "Unit 1, Lesson 01". This material should be review from Grade 11. 2. Check your answers on the website: pattersonscience.weebly.com Go to: Unit 1, Lesson 01 Answers to Homework
2	**Quantum Mechanical Model** • contributions of Planck, Einstein, de Broglie, Heisenburg and Schrodinger • difference between an orbital and an orbit • question types for quantum mechanics • aufbau Principle, Hund's Rule and Pauli Exclusion Principle	**Note**: The Quantum Mechanical Model of the Atom **Text**: p. 131 – 133, 137 – 138 and 142 – 146	1. Homework questions for Unit 1, Lesson 02 at the end of the note. 2. Check your answers on the website. Go to Unit 1, Lesson 02 Answers to Homework
3	**Electron Configurations** • full and condensed format • exceptions to the predicted electron configurations (electron promotion) **Orbital Box Diagrams**		1. Homework questions for Unit 1, Lesson 03 at the bottom of last nights homework (bottom of the homework page for Unit 1, Lesson 02.
4	**Quantum Numbers** (n, l, m_l, m_s)	**Note**: Representing Electron Configurations using Orbital Diagrams and Quantum Numbers **Text**: p. 133 – 138 and 147 – 150	1. Homework questions at the end of the note 2. Complete handout: Homework Review of the Periodic Table 3. Complete handout "Nuclear Charge and Shielding Effect" as instructed in the homework. **Bring this completed chart to our next class!**
5	**Trends on the Periodic Table** • net nuclear attraction (Z_{eff}) explains trends across a period • shielding effect explains trends down a group • know and explain the trends on the P.T. for atomic radius, first ionization energy, electronegativity, electron affinity and ionic radius **Ionization Energies** • first, second, third etc. ionization energies	**Note**: Electron Configurations and Trends on the Periodic Table **Text**: p. 152 – 157	1. Complete questions on handout "Homework on Periodic Trends" 2. Do "Unit 1, Chapter 3, Review" on the web page. There will be a quiz at the beginning of Lesson 6. Bring your questions to Lesson 5.

Unit 1, Lesson 01: Atomic Models from Dalton to Chadwick
(p. 118 – 130)

Chemistry is the study of _____________.
The Atomic Theory of Matter states that all matter is made of ___________.
The following researchers have contributed to our understanding of the composition of atoms:

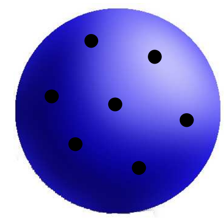
1. **John Dalton** (1766 – 1844) proposed the first "modern" atomic model:
   - matter is made of tiny solid spheres called __________
   - atoms are ____________________ and _______________________
   - each element has its own kind of atom
   - all atoms of the same element are ___________________

Billiard-Ball Model

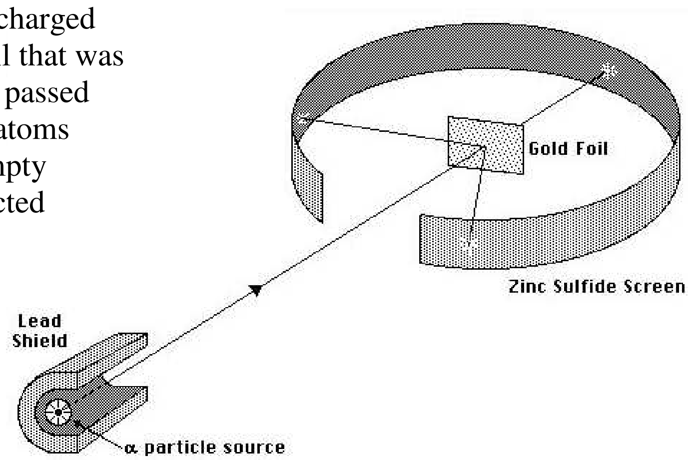


2. **J.J. Thomson** (1856 – 1940) used the cathode ray tube and found that:
   - atoms can be broken down into smaller _____________________ particles
   - he discovered the ____________, which he assigned a _____________ charge
   - atoms are _____________ overall, so he proposed that the rest of the atom is a __________________-charged solid matrix
   - electrons are "stuck" in the matrix like _____________ in a _______________

Raisin Bun Model

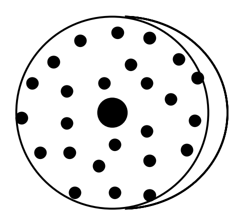


3. **Ernest Rutherford** (1871 – 1937) fired positively charged alpha particles (helium nuclei) at a piece of gold foil that was only a few atoms thick. Most of the alpha particles passed directly through the gold foil, which indicated that atoms are not solid particles. Rather, atoms are mostly empty space. A tiny number of alpha particles were deflected from the gold foil, and some bounced right back. This indicated that atoms contain a small, very dense positively charged "core" which Rutherford called the nucleus.



Rutherford:
- discovered the _______________________
- the nucleus is __________________-charged, _________ and __________
- most of the atom is ____________________
- electrons are flying ________________ around the nucleus in an _________________________

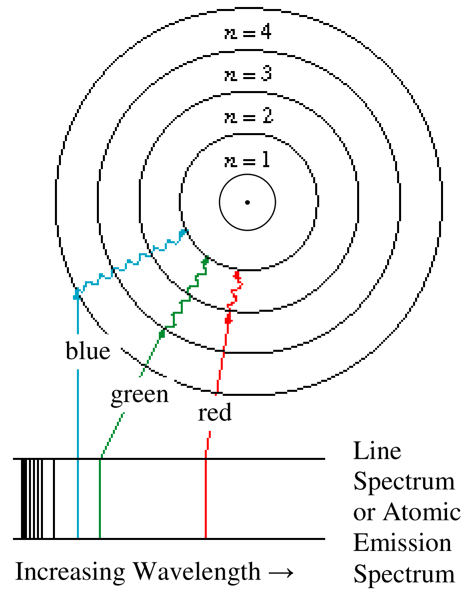
Nuclear Model with an Electron Cloud



Rutherford's model raised several questions:
- If like charges ___________, what holds the positively charged nucleus together?
- If opposite charges ______________, why don't the electrons stick to the _______________?
- The known relative masses of the atoms did not agree with the charges and masses of the nucleus calculated by Rutherford.

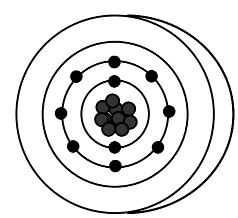
4. **Max Planck** (1858 – 1947)
   - energy comes in discrete (fixed) amounts called "____________" (singular: ________________)
   - a quantum is the smallest unit of _______________

5. **Albert Einstein** (1879-1955)
   - a quantum of energy is equivalent to a ___________________



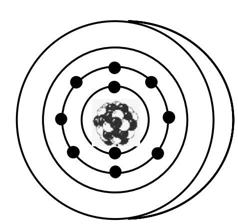
6. **Neils Bohr** (1885 – 1962)
   - he added ______________ to hydrogen gas and used a spectroscope to study the pattern of ___________ it gave off
   - in their ______________________, electrons are as close to _____________ as possible (_____________ potential energy)
   - when electrons in an atom absorb ____________, they move ____________ from the nucleus to positions of __________ potential energy
   - when electrons fall back closer to the nucleus, their potential energy _______________ and they give off this energy as ___________
   - the wavelength (____________) of the light indicates how much ______________ the electron releases
   - if electrons were found randomly in a ____________ around the nucleus, excited electrons would be all distances from the nucleus so they would produce ____________________ of light in a __________________ spectrum, like a ______________
   - Bohr found that excited electrons produce only certain colours (_______________________) of light in a pattern called a _________________________ spectrum
   - he concluded that electrons must be only ____________, _____________ distances from the nucleus
   - electrons ____________ the nucleus in discrete energy levels called "______________"
   - Bohr called the shells "________________ quantum levels" and assigned each shell an integer value: ________, _________, __________ …. _______
   - each shell can hold a maximum of _______ electrons

Bohr's Planetary Model



7. **James Chadwick** (1891 – 1974)
   - when he bombarded beryllium atoms with alpha particles, they gave off a beam of particles that was not affected by a ________________
   - he discovered the _______________, which is found in the ____________ and has ___________________
   - atoms of the same element with different numbers of neutrons are called __________________

Planetary Model with Neutrons



Unit 1, Lesson 01: Summary of Atomic Structure so far…

Atoms are made of sub-atomic particles:

- Protons: found in ______________, charge of ________, mass of _______
- Neutrons: found in _____________, _______ charge, mass of _________
- Electrons: found in ____________ around nucleus, charge of _______, mass is _______________

Atomic Number: the number of _______________ in the nucleus of an atom

- symbol is "______"
- defines the ______________ of the atom

eg. Z = 12, atom is ____________________   Z = 47, atom is _______________

Mass Number: the number of ______________ + _______________ in the nucleus of an atom

- symbol is "______"
- it is a _______________ value, it has _________________
- it is ________ reported on periodic table
- isotopes are identified by their mass number

Examples:

Isotope	**Pb – 206**	**Pb – 207**	**Pb - 208**
Atomic Number (Z)			
Mass Number (A)			
Number of Neutrons (A – Z)			

Average Atomic Mass (AAM): the ________________ average mass of all _____________ of an element

- reported on periodic table, units are _______
- AAM on periodic table is usually close to the mass number of the most ________________ isotope

eg. AAM of carbon is __________________, so most abundant isotope is probably ____________

eg. most abundant isotope of argon is probably ______________

Ions: charged atoms

- if the number of electrons equals the number of protons, the atom is _____________
- if there are more electrons than protons, the ion is __________________ charged (an ____________)
- if there are fewer electrons than protons, the ion is __________________ charged (a ____________)
- atoms tend to gain or lose electrons to obtain a ______________________ electron arrangement and become ______________________ with a ____________________

eg. Mg loses _____ electrons, forms _________ ions

Standard format:

$^{16}_{8}\mathrm{O}^{2-}$

16 → ______________________

2- → ___________

O → Chemical symbol

8 → ______________________

protons (atomic number, Z) = ______

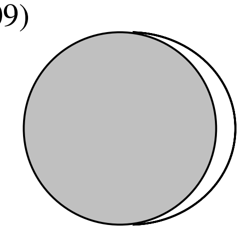
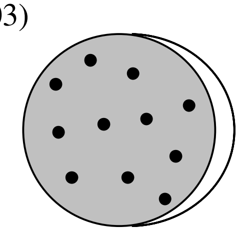
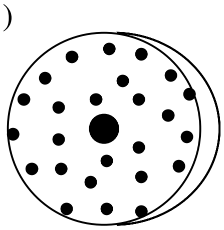
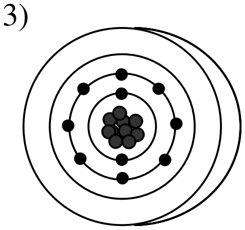
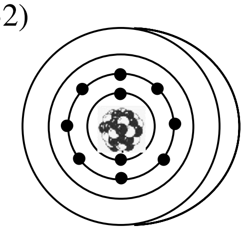
neutrons (mass number – atomic number) = ______

electrons (2 more electrons than protons) = ______

Unit 1, Lesson 01: Homework

1. Read pages 118 – 130.
2. Answer questions on page 130: 1, 2, 4, 5, 6, 7
3. Explain why Bohr's model of the atom had to be modified. (It has two fundamental flaws).
4. Complete the summary chart below. Pay specific attention to location of the electrons in the atom.

Atomic Models from Dalton to Chadwick

Researcher and Model of Atom	**Features and Limitations of this Model of the Atom (include how it is different from previous models)**
John Dalton (1809) 	**Billiard Ball Model**
J. J. Thomson (1903) 	**Raisin Bun Model**
Ernest Rutherford (1911) 	**Electron Cloud Model**
Neils Bohr (1913) 	**Planetary Model**
James Chadwick (1932) 	**Planetary Model with Neutrons**

5. Complete the following chart. Report average atomic mass to 2 decimal places.

Element	**Atomic Number (Z)**	**Number of Protons**	**Number of Electrons**	**Overall Charge**	**Number of Neutrons**	**Mass Number (A)**	**Average Atomic Mass (u)**	**Isotope (eg. Ag – 107)**
Mg				2+		24		
	47		46		62			
				2-				S - 33
			12	0	13			
			36	2-		80		
			78					Au - 197
			12		14	26		
		38		2+	50			
O			10			16		
	50		46		69			
Sb				3+		118		
			76	3+	117			
				5+				Sb-118
		78		1+		195		
			38	0		90		

6. Referring to the chart above, and using the format "Ag-107", write:

a) any atoms that are isotopes of each other

b) any atoms that are ions of each other

7. Using your Periodic Table and the format "Ag-107", predict the most common isotope for each of the following elements:

a) Mg ________________ d) H ________________ g) F _______________

b) Sr _________________ e) S _________________ h) Ar ______________

c) Al _________________ f) Na ________________ i) Pb _______________

8. Using only your Periodic Table, write the ion that will be formed (if any) by each of the following elements:

a) Mg ____________ d) Br _____________ g) Cs ____________

b) S ______________ e) N ______________ h) Ne ____________

c) Al _____________ f) Ba _____________ i) Te ____________

9. Atoms and ions that have the same number of electrons are said to be **<u>isoelectronic</u>**. List three atoms or ions, including the charge on each, that are isoelectronic with the following:

a) Ar _______________________ c) Ne _______________________ e) S^{2-} _____________________

b) Br^{1-} ____________________ d) Sr^{2+} __________________ f) Na^{1+} ___________________

Unit 1, Lesson 02: The Quantum Mechanical Model of the Atom

Recall: The number of protons (___) in an atom determines both:

- the ______________ of the atom
- the number of _________________ in a _____________ atom

> It is the _______________ and _________________________ of electrons in an atom that determines the atom's _______________ and _________________ properties.

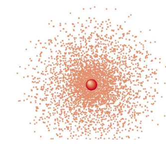
So, how are the electrons arranged?

1. **Dalton**: atoms are indivisible. There are no such things as electrons
2. **Thomson**: electrons are ____________________ in a solid, positively-charged _____________
3. **Rutherford**: electrons are found around the nucleus in a random _______________________
4. **Bohr**: electrons are found orbiting the nucleus in fixed, ________________________ (energy) levels
   - there are an ________________ number of principal quantum levels, n
   - each successive quantum level is ______________ from the nucleus and ______________ energy
5. **de Broglie (1924)**
   - all matter has ________________ properties
   - the wave-like behaviour of small objects such as _________________ is significant
   - because of their __________________ properties and _____________________________ repulsion, electrons do not move in simple, defined orbits as Bohr suggested
6. **Schrodinger (1926)**
   - combined de Broglie's wave-like properties of particles and Planck/Einstein's idea of ______________ _____________________________
   - developed a mathematical model called "__________________________" to predict the _________ of electrons within an atom
7. **Heisenberg (1927)**
   - Heisenberg's Uncertainty Principle: It is impossible to know BOTH an electron's _____________ (_______________) and ___________________ (trajectory or path)
   - an experiment designed to determine an electron's location (_____________) will change the electron's _______________ (trajectory or path)
   - an experiment designed to determine an electron's ____________ will change the electron's location (_________________)

Quantum Mechanical Model

- Schrodinger used mathematical wave functions to define _______________: regions in 3-D space where an electron can be found _______ of the time
- an orbital holds a maximum of ____ electrons
- principal quantum levels can be divided into ____________________ containing different types of orbitals (_______________ etc) depending on the _____________ (______________) of the electrons

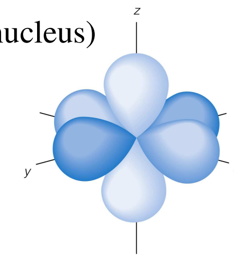
Unit 1, Lesson 02: Summary of the Quantum Mechanical Model



an "*s*" orbital

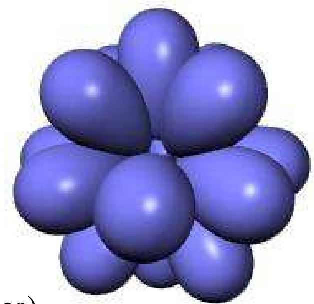
For **n = 1** (the ________ quantum level, or the energy level closest to the nucleus)
- holds a maximum of $2n^2$ electrons, or (   ) _____ electrons
- there are n^2 (   ) or _____ orbital (each orbital can hold 2 electrons)
- there is _____ (n) type of orbital:
  ______ spherical "s" orbital, called ______

For **n = 2** (the ____________ quantum level, or the second energy level away from the nucleus)
- holds a maximum of $2n^2$ electrons, or (   ) _____ electrons
- there are n^2 (   ) or _____ orbitals (each orbital can hold 2 electrons)
- there are _____ (n) types of orbitals:
  ________ spherical "s" orbital, called ______
  ________ perpendicular "p" orbitals called ______, ______ and ______
  ("p" orbitals are _____________ shaped and found at right angles to each other in three three planes)



three "*p*" orbitals

For **n = 3** (the _________ quantum level)
- holds a maximum of $2n^2$ electrons, or (   ) _____ electrons
- there are n^2 (   ) or _____ orbitals (each orbital can hold 2 electrons)
- there are _____ (n) types of orbitals
  ________ spherical "s" orbital, called ______
  ________ perpendicular "p" orbitals called ______
  ________ diffuse "d" orbitals called ______
  ("d" orbitals are large, _____________________ shaped and found in five planes)



five "*d*" orbitals

For **n = 4** (the fourth quantum level, or the fourth energy level away from the nucleus)
- holds a maximum of $2n^2$ electrons, or (   ) _____ electrons
- there are n^2 (   ) or _____ orbitals (each orbital can hold 2 electrons)
- there are _____ (n) types of orbitals:
  ________ spherical "s" orbital, called ______
  ________ perpendicular "p" orbitals called ______
  ________ diffuse "d" orbitals called ______
  ________ fundamental "f" orbitals called ______ (the "f" orbitals are large, shape unknown)

There are an _______________ number of quantum levels; each one is further and further from the nucleus. The types of orbitals are: _______________________ etc. The g, h, i etc. orbitals are for electrons in their __________________________.

Question Types:

1. The maximum number of electrons that can be held in each quantum level is ______.
2. The total number of orbitals in each quantum level is _______.
3. The number of <u>types</u> of orbitals (sub-levels) in each quantum level is _______.
4. The types of orbitals are identified with letters . . . s, p, d, f, ...

   If these occur in a given energy level there is(are) always:

   ______ s-orbitals, ______ p-orbitals, ______ d-orbitals, ______ f-orbitals
5. The maximum number of electrons which may be designated (named):

1s: ____ , 2s: ____ , 2p: ____ , 3s: ____ , 3p: ____ , 3d: ____ , 4s: ____, 4p: ____ , 4d: ____ , 4f: ____

Unit 1, Lesson 02: Homework on Quantum Mechanics

1. Read pages 131 – 133 (not Quantum Numbers, yet), 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.
3. Explain how the quantum mechanical model of the atom differs from Bohr's model of the atom.
4. Explain how an orbital is different from an orbit. Be specific.
5. How many electrons (maximum) are in quantum level 4? __________ , When n=3 ________
6. How many electrons can be designated as 3d? _________ , 4s __________, 5f ____________
7. How many types of orbitals are there in quantum level 3? _________, When n=4 ________
8. How many orbitals are there in quantum level 2? ________, When n=5 ________, n=3 ________
9. How many electrons can be held in quantum level 5? _________, When n=1 ________
10. How many orbitals are there in quantum level 1? _________ ,When n=4 ________
11. How many types of orbitals are there in quantum level 5? _______, When n=2 _______ , n=6_______
12. How many electrons can be designated 2p? _______, 4p ________, 5s ________, 6f ________
13. Circle the orbitals which **do not** exist: 3f 2p 5s 2d 4f 1d 5p 3d 1p 3s 4d 4s
14. Questions 6, 7 on page 145.
15. Questions 10, 11, 12 and 13 on page 150.

Unit 1, Lesson 03: Homework on Electron Configurations

1. Write the predicted and actual (experimentally determined) electron configurations for Mo, Ag and Au.
2. If valence electrons are found in the order that we would predict ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$ etc.), then the electrons are in their ground state and as close to the nucleus as possible. If the electron configuration is "out of order", it means that electrons are not in their ground state. Instead, they are in energy levels further from the nucleus than expected, so these electrons are in an excited state.

   Do the following electron configurations show electrons in their ground state or an excited state?

a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$ ____________________

b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$ ____________________

c) $1s^2\ 2s^2\ 2p^6\ 3p^2$ ____________________

d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 4p^2$ ____________________

e) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 6s^1$ ____________________

f) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$ ____________________

g) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 4d^4$ ____________________

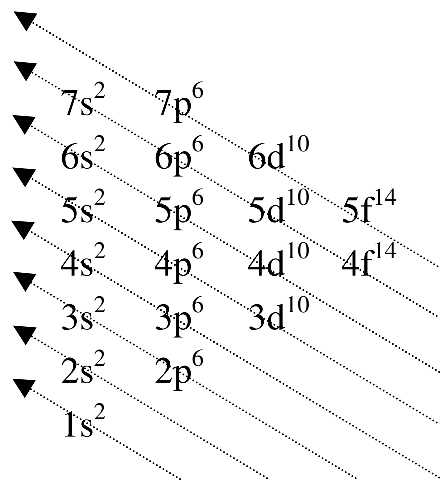
Unit 1, Lesson 03: Electron Configurations

There are three rules when writing electron configurations:

1. **Aufbau Principle**: electrons fill the ____________ available energy level (get as _________ to the _____________ as possible)
2. **Pauli Exclusion Principle**: each orbital holds a maximum of _____ electrons with opposite ______
3. **Hund's Rule**: electrons do not ________________ in an orbital until all orbitals of the same sub-level are ______________________

Electron configurations are written using the form: $1s^1$

The order of filling puts the electrons as close to the nucleus as possible. The order of filling can be read from the ____________ ____________ or remembered using the mnemonic:



$_{15}P$ ____

$_{40}Zr$ ____

$_{54}Xe$ ____

$_{76}Os$ ____

Electron Configurations using the Condensed Format

An atom's full electron shells are represented by the symbol of the nearest __________________ Noble Gas, in ____________ brackets, followed by the electron configurations for the ___________ electrons:

$_{15}P$ _________________________________ $_{40}Zr$ ________________________________

$_{54}Xe$ ________________________________ $_{76}Os$ ________________________________

Exceptions to the Predicted Electron Configurations

Chromium and molybdenum:

- their predicted configurations end in _________, but their actual configurations end in ________
- Why? it is lower energy (_____________________) to have the "d" orbitals all ________________, so one "____" electron is _________________ to the "____" sub-level

Copper, silver and gold:

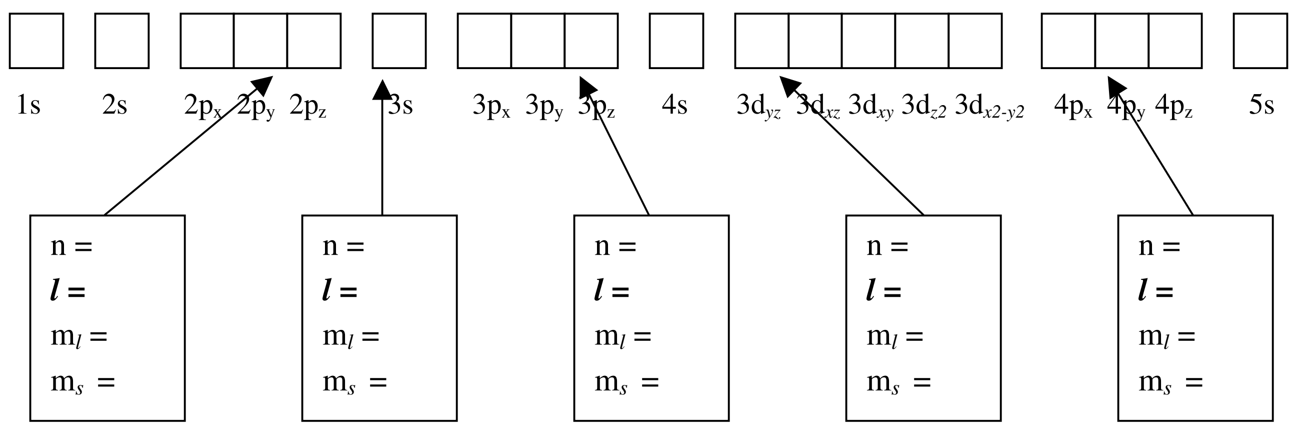
- their predicted configurations end in ________, but their actual configurations end in ________
- Why? it is lower energy (_____________________) to have the "d" orbitals all ________________, so one "____" electron is _________________ to the "____" sub-level

Unit 1, Lesson 04: Summary of Quantum Numbers

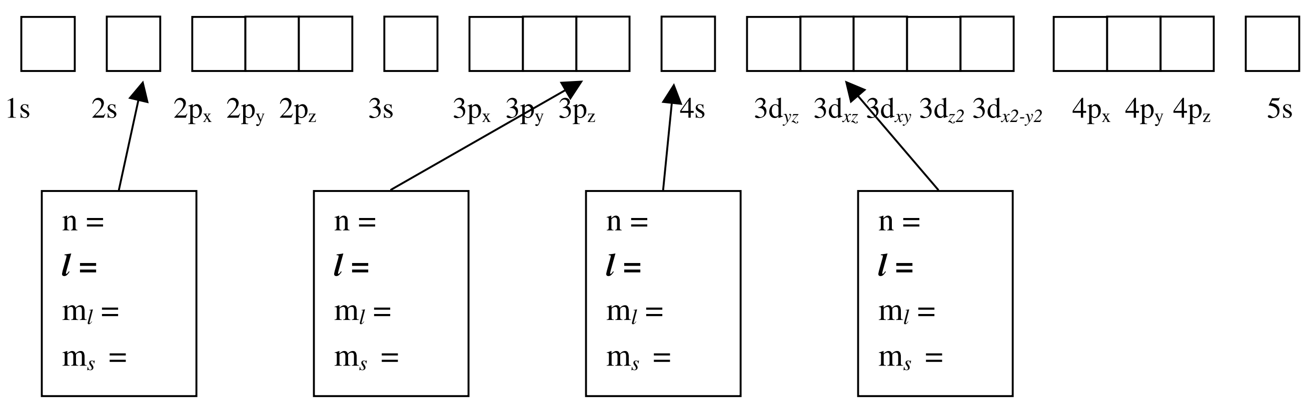
Summary: The "allowed" values for quantum numbers for each principal quantum level "*n*":

n	***l***	m_l	m_s	corresponding sub-level	number of orbitals in this sub-level
n = 1	0	0	+ ½ , - ½	1s	1
n = 2	0	0	+ ½ , - ½	2s	1
	1	–1, 0, +1		2p	3
n = 3	0	0	+ ½ , - ½	3s	1
	1	–1, 0, +1		3p	3
	2	–2, –1, 0, +1, +2		3d	5
n = 4	0	0	+ ½ , - ½	4s	1
	1	–1, 0, +1		4p	3
	2	–2, –1, 0, +1, +2		4d	5
	3	–3, –2, –1, 0, +1, +2, +3		4f	7

eg. $_{33}$As: ____ or ____________________



$_{29}$Cu ____ or ____________________



Unit 1, Lesson 04: Homework on Quantum Numbers

1. Write the quantum numbers that represent the following electrons:

a) a $5p^3$ electron would be given the quantum numbers: n = ____, l = ____, m_l= ____ and m_s =____

b) a $3s^2$ electron would be given the quantum numbers: n = ____, l = ____, m_l= ____ and m_s =____

c) a $4f^6$ electron would be given the quantum numbers: n = ____, l = ____, m_l= ____ and m_s =____

2. What are the allowable (possible) values for l when:

a) n = 4: ___________________

c) n = 1: ___________________

b) n = 3: ___________________

d) n = 5: ___________________

3. What are the allowable (possible) values for m_l when:

a) n = 4, l = 3: ____

b) n = 3, l = 1: ____

c) c) n = 2, l = 0: _______________________________________

d) d) n = 5, l = 4: _______________________________________

4. Write the principal quantum number and ***letter*** indicating orbital shape for each of the following:

a) n = 2, l = 1________

c) n = 4, l = 3 ________

e) n = 4, l = 1 ________

b) n = 3, l = 2________

d) n = 1, l = 0 ________

f) n = 2, l = 0 ________

5. State whether the following sets of quantum numbers are possible ( ) or impossible (X). Identify the values which are incorrect or impossible, if any.

a) n = 3, l = 3, m_l= -1 and m_s = + ½ ____

b) n = 5, l = 2, m_l= -1 and m_s = - ½ ____

c) n = 2, l = 0, m_l= 0 and m_s = - ½ ____

d) n = 3, l = 1, m_l= 0 and m_s = 0 ____

e) n = 1, l = 0, m_l= +1 and m_s = + ½ ____

f) n = 0, l = 0, m_l= 0 and m_s = + ½ ____

g) n = 4, l = 1, m_l= +1 and m_s = + ½ ____

h) n = 2, l = 1, m_l= -2 and m_s = - ½ ____

Unit 1, Lesson 04: Homework Review of Periodic Table

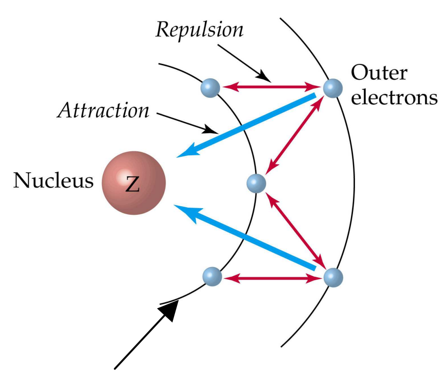
1. Read pages 133 – 138.
2. On page 136, answer questions 1 – 5.
3. On page 138, answer questions 2, 3, 5, 6.
4. Read pages 147 – 150.
5. To see how electron configurations are related to an element's position on the periodic table, write the name of the last valence electron of each element (eg. $3d^5$) in the appropriate square of the Periodic Table below. Use the ***predicted*** electron configurations for Cr, Mo, Cu, Ag and Au.
6. On the Periodic Table below, label the:

a) Group numbers and Periods
b) s,p,d and f blocks of elements
c) the transition elements and inner-transition elements
d) Noble gases, Alkali metals, Alkaline Earth metals, and Halogens

$1s^1$	$1s^2$																
$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$			

7. On the page "Nuclear Charge and the Shielding Effect: Explaining the Trends on the Periodic Table" (handed out in class), for each element complete the:

a) electron configuration
b) Rutherford-Bohr diagram
c) Nuclear charge (the number of protons in the nucleus = atomic number = Z)
d) Shielding effect (the number of electrons in the full shells between the nucleus and the valence shell)
e) Net Nuclear attraction (the nuclear charge subtract the shielding effect). Net nuclear attraction is the effective (Z_{eff}) or actual attraction that exists between a valence electron and the nucleus.
f) Use the numbers on the back of your Periodic Table to complete the ionization energy (First Ionization Potential, V), electronegativity and Atomic Radius (Δ, Angstroms)



Shielding Effect

Bring the completed sheet to class for our next lesson!

Unit 1, Lesson 05: Trends on the Periodic Table

Refer to chart “Nuclear Charge and the Shielding Effect: Explaining the Trends on the Periodic Table”.

__Shielding effect__ **(SE)** is defined as the number of _______________ in the _______________________ between the _____________ and the outer _______________________

- as the shielding effect increases, the valence electrons are _____________ from the _____________, so the attraction between the nucleus and the valence electrons _______________
- across each period (row →), the shielding effect _________________
- down each group (column), the shielding effect _________________, so the attraction between the nucleus and electrons _______________

__Net nuclear attraction__ (aka. ____, __________) is defined as the _________________________ (___) subtract the __________________________

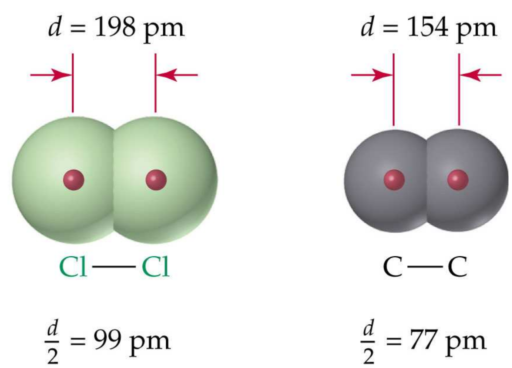
- Z_{eff} represents the _____________________________________ between the _____________ and the _____________ electrons
- as Z_{eff} increases, the valence electrons are pulled ___________ and ___________ to the ___________
- across each period (row →), Z_{eff} ______________
- down each group (column), Z_{eff} is _______________

Trends across a period (→) are explained by ____
Trends down a group (↓) are explained by ____

1. Atomic Radius: one half the distance between the ______________ of two of the same type of atom, _____________ bonded together or as a ____________ under controlled conditions

a) down a group (↓), atomic radius _______________ because ____________________________

b) across a period (→), atomic radius ______________ because ____________________________ which pulls the valence electrons _______________ to the nucleus



2. First Ionization Energy (   ): the amount of energy required to remove the _________________ electron from a ________________________, _______________, ________________ atom

a) down a group (↓), IE_1 _______________ because _____________________________ so the valence electrons are _________________ from the nucleus and ______________ to remove

b) across a period (→),IE_1 ______________ because _____________________________ which holds the valence electrons ________________________ to the nucleus

3. Electronegativity (   ): a measure of the relative ________________ that an atom has for the ______________ in a ___________, compared to _______________

a) down a group (↓), EN _______________ because _____________________________. The valence electrons are _________________ from the nucleus and ______________ attracted to it

b) across a period (→), EN ______________ because _____________________________. The valence electrons are ______________________________________ to the nucleus

4. Electron Affinity ( ): is the change in ________________ that occurs when an electron is __________ to a ________________________, ________________, _________________ atom

a) metals and Noble gases: do _________________ another electron

- energy must be ____________ to make the electron "stick" so EA is ____________________
- down a group (↓), EA ________________ (moves closer to __________) because __________________ and the _________________ for a new electron is ____________

b) non-metals: _________ another electron

- energy is ________________ when an electron is added so EA is ________________
- across a period (→), EA for non-metals _______________ (becomes ______________ and more _____________) because ______________________________ and the elements are closer to a ________________________

5. Ionic Radius: the radius of an ion

a) metal atoms ____________ electrons when they form ions

- they have ____________ electrons, so metal ions are ______________ than their parent atom

b) non-metal atoms ____________ electrons when they form ions

- they have __________ electrons, so non-metal ions are ______________ than their parent atom

The ion with the _________ protons and _____________ electrons will have the smallest radius.

eg. Put the following atoms and ions in order of ***increasing*** radius: P P^{3+} P^{5+} P^{3-}

6. Successive Ionization Energies

- multi-electron atoms can have many ionization energies as more and more electrons are __________

$_{11}Na$ $1s^22s^22p^63s^1$ → ___________________________________ $IE_1 =$ 5.1 eV

$_{11}Na^{1+}$ $1s^22s^22p^6$ → ___________________________________ $IE_2 =$ 47.3 eV

$_{11}Na^{2+}$ $1s^22s^22p^5$ → ___________________________________ $IE_3 =$ 71.7 eV

$_{11}Na^{3+}$ $1s^22s^22p^4$ → ___________________________________ $IE_4 =$ 98.8 eV

Notice:

- the IE for each successive ionization ________________ because we are removing an electron from a more and more ____________________________ ion
- there is a huge jump in energy to remove a ______ electron (break a _________________________)
- similarly, it also takes more energy to remove ______, ______, ______ and ______ electrons (any time you are breaking up a _________ or ____________________ energy sub-level)

Example: from the ionization energies for an unknown element, identify its group number:

$IE_1 =$ 8.3 eV

$IE_2 =$ 25.1 eV

$IE_3 =$ 37.9 eV

$IE_4 =$ 259.3 eV

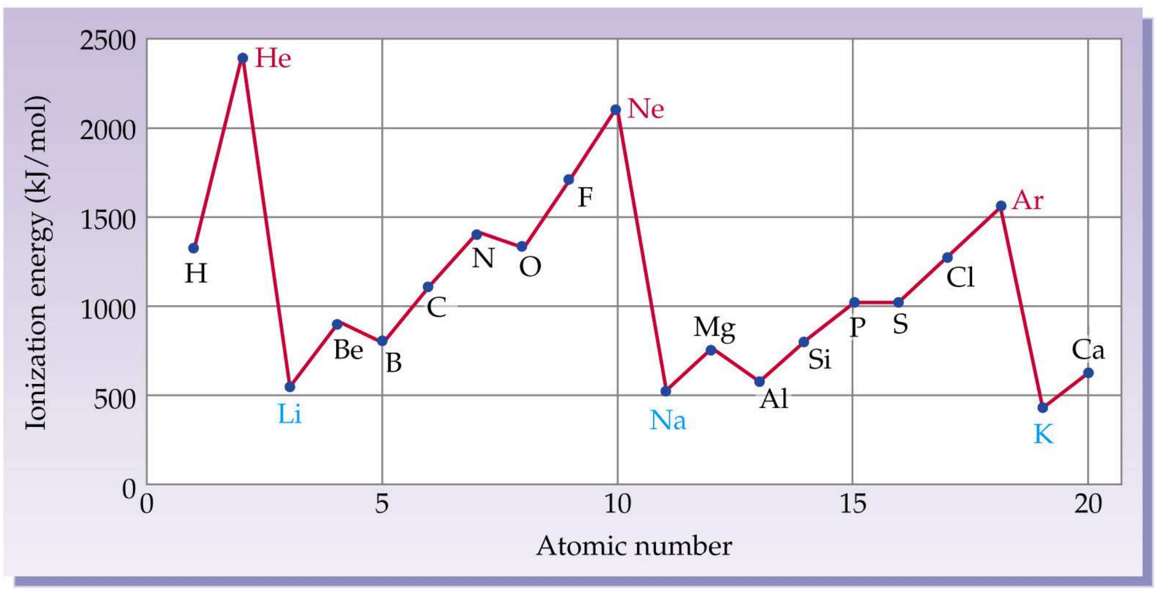
$IE_5 =$ 340.1 eV

Unit 1, Lesson 05: Homework on Periodic Trends

1. Read pages 152 - 157 in your text.
2. An atom of argon has 18 electrons while an atom of sodium has only 11 electrons. However, an atom of argon is SMALLER than an atom than sodium. Explain why this is true.
3. What three factors influence the magnitude of ionization energy?
4. Explain why lithium has a lower first ionization energy than fluorine.
5. Explain why potassium has a lower first ionization energy than lithium.
6. As a general rule, positive metal ions are smaller than their neutral metal atoms, while negative non-metal ions are larger than their neutral non-metal atoms. Explain these patterns.
7. Put the following in order of ***increasing*** radius. Explain why you put them in this order:

   a) Mn, Mn^{2+}, Mn^{4+} b) P^{3+}, P^{3-}, P, P^{5+}

First Ionization Energy vs. Atomic Number for the First 20 Elements



8. Refer to the graph above to answer the following questions:

a) What is significant about the IE_1 for the Noble gas elements? Explain.

b) What is significant about the IE_1 for the Alkali metal elements? Explain.

c) How do the IE_1 for the Period 2 elements compare to those of Period 3? Explain.

d) The IE to remove a $2s^2$ electron is higher than the IE to remove a $2p^1$ electron. Explain.

e) The IE to remove a $2p^3$ electron is higher than the IE to remove a $2p^4$ electron. Explain.

9. The first eight ionization energies for phosphorus are shown in the table below:

ELEMENT	IONIZATION ENERGIES (eV, Electron Volts)							
	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8
Phosphorus	11.0	19.7	30.2	51.4	65.0	220.4	263.3	309.3

a) Write the first six ionization reactions for phosphorus. Include the ionization energy required for each reaction.

b) Explain why IE_6 is so much higher than IE_5.

10. The successive ionization energies for four unknown elements are reported below. Prepare a graph showing ionization energy vs. successive ionization. (You can do one single graph or four separate graphs). For each element:

a) Identify which ionization reaction is breaking a stable octet electron arrangement.
b) How many valence electrons does each neutral atom have?
c) Identify which group on the Periodic Table each element is found in.

ELEMENT	IONIZATION ENERGIES (eV, Electron Volts)							
	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8
Unknown A	13.6	35.1	54.9	77.4	113.9	138.1	739.1	871.1
Unknown B	13.0	23.8	39.9	53.5	67.8	96.7	114.3	348.3
Unknown C	11.0	19.7	30.2	51.4	65.0	220.4	263.3	309.3
Unknown D	6.0	18.8	28.4	120.0	153.8	190.4	241.9	285.1

11. Answer questions 3, 4, 5, 6b,c,g,h, 7 and 8 on pages 157- 158.

12. Explain the difference between electron affinity and electronegativity.

13. An atom has a large negative value for electron affinity. What type of atom would this be?

14. An atom has a large positive value for electron affinity. What type atom of would this be?

15. Summarize the trends on the Periodic Table for electronegativity, first ionization energy and atomic radius in a way that is meaningful to you. Be able to explain and apply each of these trends.