SCH 3UI Unit 2 Outline Up to Unit Test
Atomic Theory and the Periodic Table

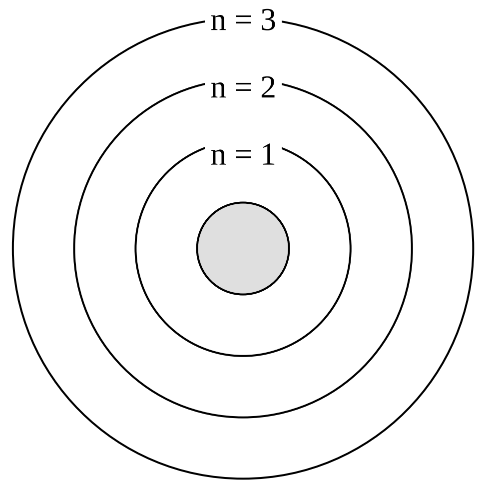
Lesson	Topics Covered	Homework Questions and Assignments
10	**Note: Introduction to the Quantum Mechanical Theory of the Atom** • Review of Bohr's Atomic Theory • Heisenberg Uncertainty Principle • Orbitals (demo)	• Re-read the note: The Quantum Mechanical View of Atoms • Work on articles and questions about radio-isotopes
11	**Note: Electron Arrangement in Quantum Theory** • principle quantum number (n) • shapes of orbitals • summary of question types for Quantum theory	• Answer questions on Handout: Practice Questions: The Quantum Mechanical Model of the Atom • Be able to recognize what is being asked in each type of question
12	**Note: Electron Configurations** • aufbau principle • Pauli exclusion principle • Hund's Rule	• Use the energy level diagram on the handout "Ways to Remember the Order of Filling" to write electron configurations for elements: 10-20, 25, 34, 40, 50, 55, 60, 78, 85 and 92
13	**Electron Configurations and the Periodic Table** • Chemical groups (columns) • Chemical periods (rows) • "Reading" electron configurations	• study the patterns in electron configurations in the Periodic Table • Begin Take Home Quiz: Electron Configurations (due at the beginning of the 15th lesson)
14	**Note: Ionization Reactions and the Periodic Table** • Groups on the Periodic Table • ionization reactions • "isoelectronic" atoms and ions	• Complete Practice Questions: Electron Configurations and the Periodic Table • Complete Take Home Quiz: Electron Configurations (due at the beginning of the next lesson)
15	**Lab #3: Trends on the Periodic Table** • prelab notes and demos	• Begin Lab report #3. You can answer any questions involving ionization reactions
16	**Lab #3: Trends on the Periodic Table** • complete lab, take up errors	• Work on lab report #3 (due at the beginning of lesson 18)
17	Background for explaining the trends on the Periodic Table • nuclear charge • shielding Effect • net nuclear attraction (NNA), which is also known as effective nuclear charge or Z_{eff}) **Note: Explaining the Trends on the Periodic Table**	• Complete the Handout: Nuclear Charge and the Shielding Effect: Explaining the Trends on the Periodic Table • Understand the meaning of shielding effect and net nuclear attraction • Use shielding effect and NNA to explain the trends on the PT
18	**Note: Explaining the Trends on the Periodic Table (cont)** • reactivity of metals (metallic character) • reactivity of non-metals • atomic radius • first ionization energy • electronegativity	• Complete: Practice Questions: Trends on the Periodic Table • Begin Unit 2 Review • Unit test: ____________________

The Quantum Mechanical View of Atoms

We now know that it is the arrangement of electrons around the nucleus of an atom that determines the chemical and physical properties of that atom.

Based on his observations of the bright line spectra of hydrogen atoms, Bohr developed the "Planetary" model of the atom with the following characteristics:

- electrons are found in the space around the nucleus
- electrons occupy discrete energy levels, called quantum levels, around the nucleus
- electrons have the lowest energy when they are closest to the nucleus (in their ground state)
- electrons tend to move toward positions of lowest energy
- when electrons have energy added to them (are excited), they jump to a higher energy level which is further from the nucleus. Eventually, the electrons "fall back" to a lower energy level, and release the added energy as light
- depending on which quantum levels an electron is moving between, it will give off light of different wavelengths (different colours)
- electrons travel in precise and defined paths called orbits



Bohr based his model of the atom on the study of hydrogen. This was an excellent atom to study because hydrogen atoms have only one proton and one electron, so it was "easy" to discover a pattern. Unfortunately, Bohr's atomic model is only really accurate for hydrogen and helium. The arrangement of electrons predicted by Bohr's model did not agree with actual observed trends in the chemical and physical properties of more complex elements. A final (?) revision of the model was required.

When atoms have more than just a few electrons, the electrons are being attracted toward the nucleus at the same time that they are being repelled by each other. Because of the electron-electron repulsion and the wave-like behaviour of electrons, electrons **DO NOT** travel in precise, defined, predictable orbits as Bohr suggested. That is, we can never really know exactly where an electron is going to be.

In 1927, Heisenberg expressed this idea as the Uncertainty Principle: it is impossible to know exactly the location and orbit (trajectory) of an electron within an atom. A new model, called the Quantum Mechanical Model, was needed.

The Quantum Mechanical Model of the atom proposes that electron positions should be described in terms of the regions around the nucleus where an electron is most likely to be found. The three dimensional regions in the space around the nucleus where there is a high chance of finding an electron (over 90% probability) are called "**orbitals**". Each orbital can hold only two electrons.

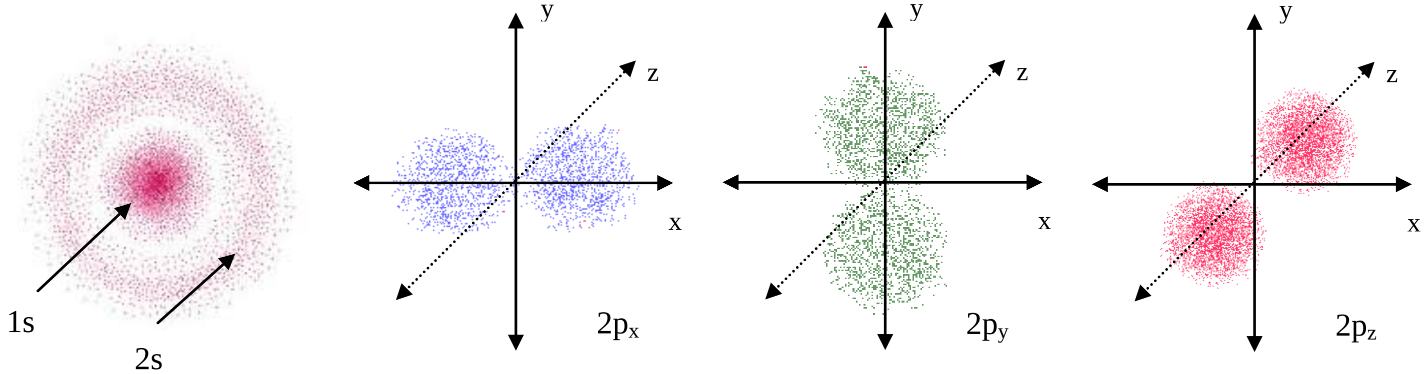
As in the Bohr model of the atom, electrons are arranged in energy levels which are identified by an integer, n, called the principal quantum number. In quantum mechanics, however, each main energy level (n) is sub-divided into a number of energy sublevels, each containing a different type of orbital. The number of sublevels (types of orbitals) increases as the value of "n" increases.

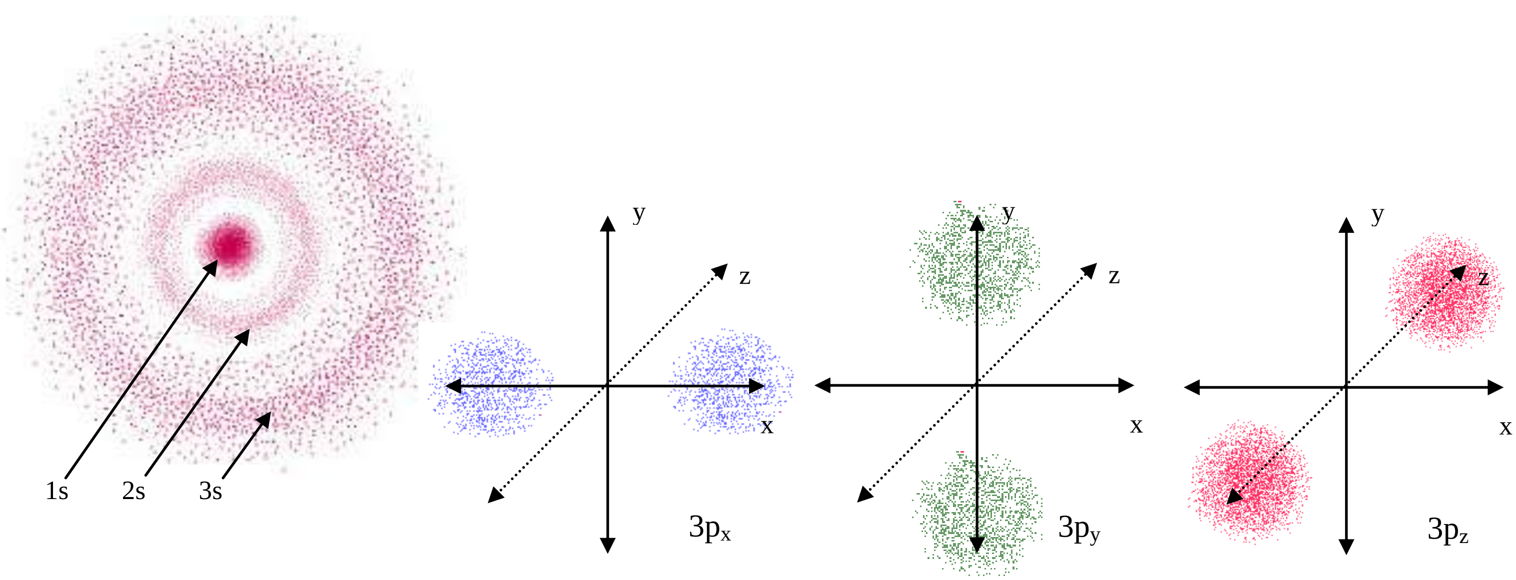
Shapes of Orbitals

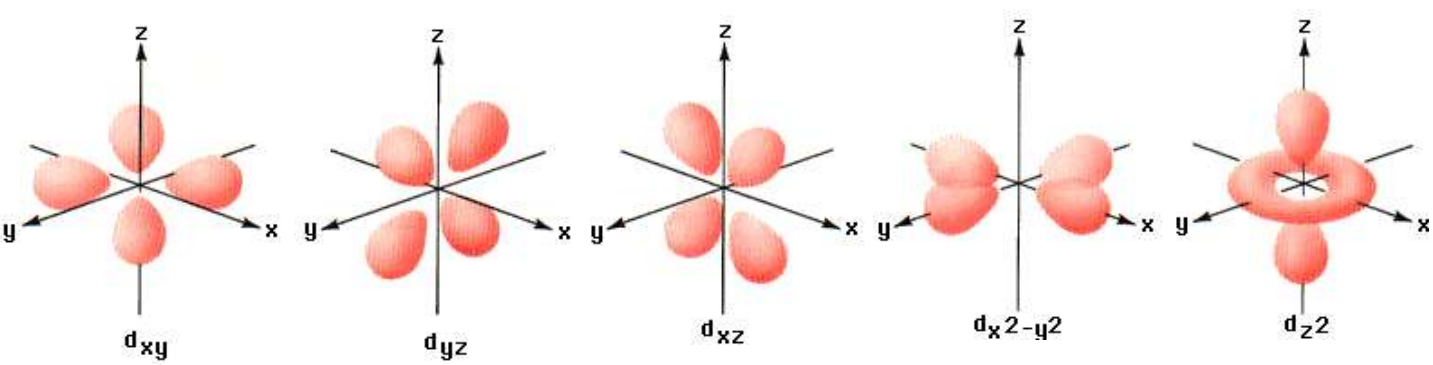
Electrons do not travel in fixed, predictable orbits. Rather, we must describe an electron's location in terms of the probability of finding an electron in a certain location. An **orbital** is defined as a region in three-dimensional space where there is greater than 90% probability of finding an electron. An orbital can hold two electrons.

We describe orbitals in terms of:
1. their principal quantum number (n = 1, 2, 3…) which tells us how far the orbital is from the nucleus
2. their shape (s, p, d, f, g, h, i etc.)
3. their orientation in 3-D space (in the x, y and z planes)

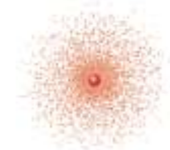
The shapes of some orbitals are shown here:







Summary of the Quantum Mechanical Model

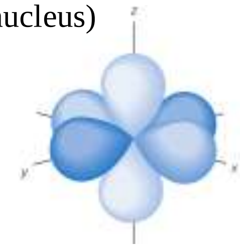


an "s" orbital

For n = 1 (the ________ quantum level, or the energy level closest to the nucleus)

- holds a maximum of **$2n^2$** electrons: ________________
- there are **n^2** orbitals: ________________ (each orbital can hold 2 electrons)
- there are **n *types*** of orbitals:

  ______ spherical "s" orbital, called ______



three "*p*" orbitals

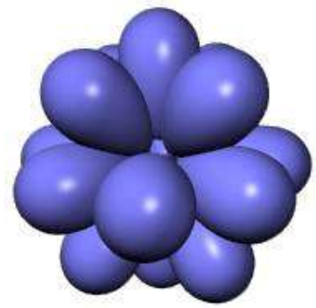
For n = 2 (the _____________ quantum level, or the second energy level away from the nucleus)

- holds a maximum of **$2n^2$** electrons: ________________
- there are **n^2** 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 planes)



five "*d*" orbitals

For n = 3 (the __________ quantum level; the third energy level away from the nucleus)

- holds a maximum of **$2n^2$** electrons: ________________
- there are **n^2** 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)

For n = 4 (the fourth quantum level, or the fourth energy level away from the nucleus)

- holds a maximum of **$2n^2$** electrons: ________________
- there are **n^2** 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: ____

Practice Questions: The Quantum Mechanical Model of the Atom

1. Re-read the handout "The Quantum Mechanical View of Atoms" and answer the following questions:

a) Why is the electron arrangement in an atom so important?
b) Which characteristic of Bohr's model of the atom is now known to be incorrect?
c) Give two reasons why electrons do not travel in precise, defined orbits.
d) State Heisenberg's Uncertainty Principle. What does it mean?
e) Define "orbital". How is an orbital different than an orbit?
f) How is the quantum mechanical model of the atom the same as Bohr's model? How is it different?
g) How many electrons can an orbital hold?
h) What is "n"?

2. Re-read the handout "Summary of the Quantum Mechanical Model" and answer the following questions:

a) What is the **maximum number of electrons** in the quantum level when:

n = 1 _____ n = 2 _____ n = 3 _____ n = 4 _____ n = 5 _____ In general? _____

b) What is the **total number of orbitals** in the quantum level when:

n = 1 _____ n = 2 _____ n = 3 _____ n = 4 _____ n = 5 _____ In general? _____

c) How many <u>**types**</u> **of orbitals** are there in the quantum level when:

n = 1 _____ n = 2 _____ n = 3 _____ n = 4 _____ n = 5 _____ In general? _____

d) How many:

i) s-orbitals are there in the second principal energy level: ______
ii) p-orbitals are there in the second principal energy level: ______
iii) d-orbitals are there when n = 3: ______
iv) s-orbitals are there when n = 4: ______
v) f-orbitals are there when n = 5: ______

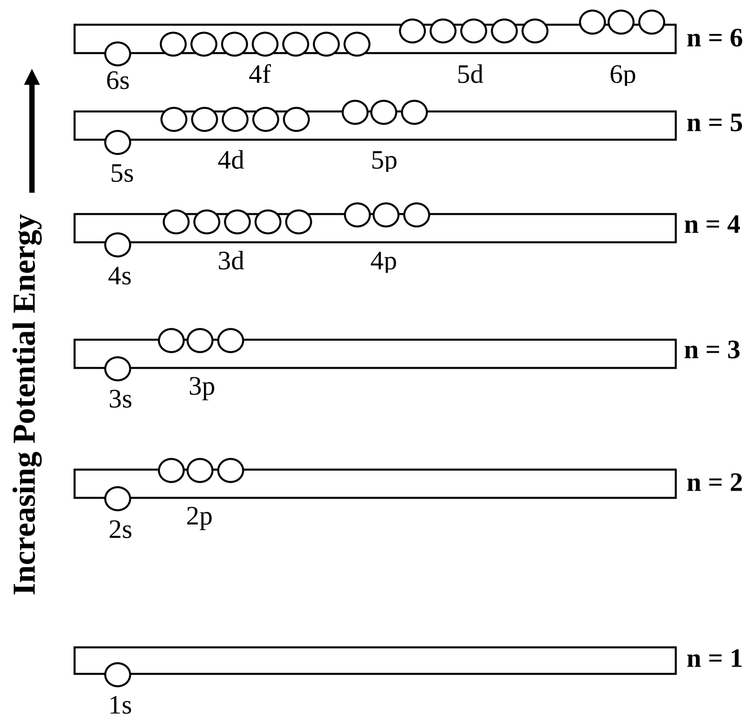
e) How many electrons can be designated (named):

3p ______ 4d ______ 5f _______ 3d ______
3s ______ 1s ______ 2p _______ 4f ______

3. **Sample Test Questions**:

a) What is the maximum number of electrons in quantum level 4? ______ When n=3? ______ n= 5 _____
b) How many electrons can be designated as 3d ______ , 4s ______, 5f ______, 2p ______
c) How many types of orbitals are there in quantum level 3? ______ When n=2? _______ n = 5 _______
d) How many orbitals are there in quantum level 2? _______ When n=5? _______ n=3 _______
e) How many electrons can be held in quantum level 5? ______ When n=1? _______ n=4 _______
f) How many orbitals are there in quantum level 1? ______ When n=4? _______ n=3 _______
g) How many types of orbitals are there in quantum level 5? _______ When n=2? _______ , n=6_______
h) How many electrons can be designated 2p _______, 4p ________, 5s ________, 6f ________
i) Circle the orbitals which **do not** exist: 3f 2p 5s 2d 4f 1d 5p 3d 1p 3s 4d 4s

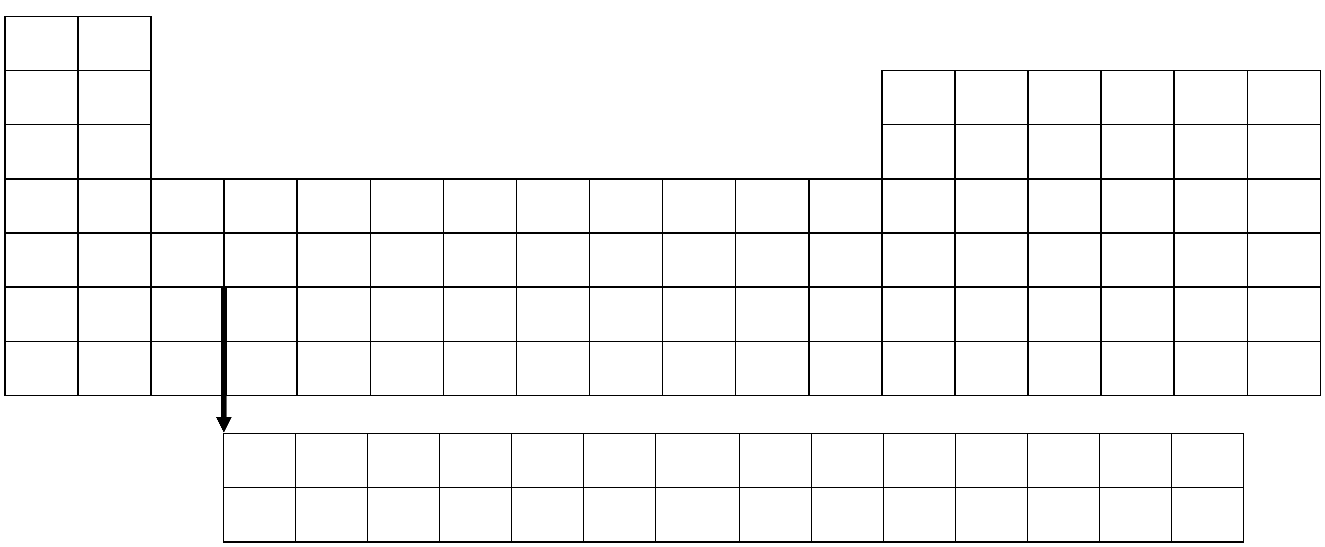
Ways to Remember the Order of Filling for Electron Orbitals



1. Energy Level Orbital Diagrams can be used to remember the order of filling for electron orbitals.

2. The order of filling for the electron orbitals can also be found using the following mnemonic:

3. The Periodic Table can be used to know the order of filling for electron orbitals.



Electron Configurations and the Periodic Table

All elements in the same **column** on the Periodic Table (called a ________ or _________) have the same number of ________________________, so they have similar ___________ and ___________ properties.

All elements in the same **row** on the Periodic Table (called a __________) have the same number of __________________________________ (   ).

You can "read" the electron configuration of an element to determine the Group and Period in which an element is located, and from this you can determine the identity of the element.

Reading Electron Configurations

- draw a line in front of the last "s" term
- the number in front of the last "s" term tells us the Period (row) of the element
- the sum of the number of electrons after the line (these electrons are in the last energy level so they are the valence electrons), is the group number for the element

Example 1: Identify the element which has the following electron configuration:

$$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,|\,4s^2$$

this element has _____ valence electrons so it is found in

this element is found in the

□the element is _______________

Example 2: Identify the element which has the following electron configuration:

$$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,|\,5s^2\,4d^{10}\,5p^5$$

this element is found in the

this element has (___ + ____ + ___) = ____ valence electrons, so it is found in _____ ________.

□the element is _______________

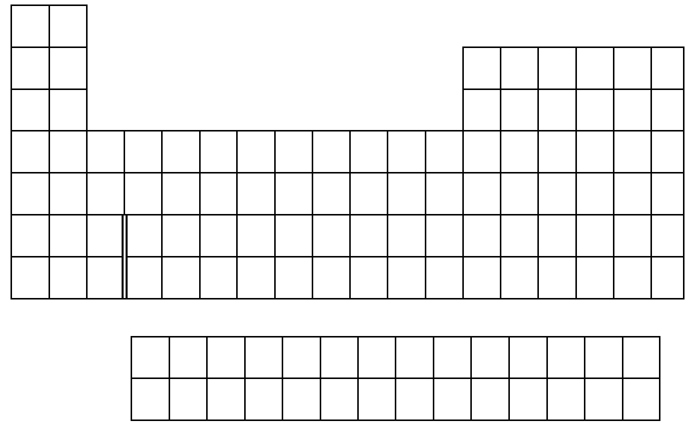
Example 3: Identify the element which has the following electron configuration:

$$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^1$$

□the element is _______________

Practice Questions: Electron Configurations and the Periodic Table

1. On the sketch of the Periodic Table, label:
a) the periods
b) the group numbers
c) colour the "s block" elements red
d) colour the "p block" elements blue
e) colour the "d block" elements green
f) colour the "f block" elements yellow



2. "Read" the following electron configurations to determine the identities of the following elements:
a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$ ________________________
b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$ ________________________
c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6$ ________________________
d) $1s^2\ 2s^2\ 2p^2$ ________________________
e) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$ ________________________

3. Write ionization reactions for the following elements:
a) N

b) Sc

c) P

d) Rb

e) Se

4. Explain what is meant by the term "isoelectronic".

5. What are four ions, with their charges, that are isoelectronic with each of the following:
a) neon ____________________________
b) argon ____________________________
c) krypton ______________________________
d) a S^{2-} ion ______________________________

6. A calcium ion, Ca^{2+}, is isoelectronic with argon. Does this mean that calcium has turned into argon? Explain.

7. Write the electron configuration for the element with atomic number 117. In which chemical group does it belong? Predict the charge on the ion that it will form.

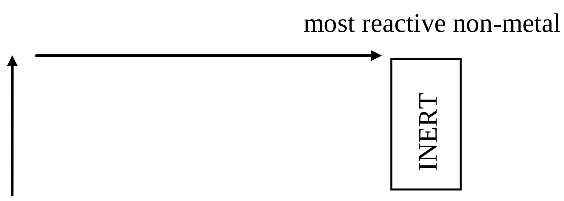
Visual Summary of the Trends on the Periodic Table

Many properties of atoms show trends or patterns as we move across a period or up and down within a group of elements within the periodic table. Some of these "periodic properties" are: reactivity, metallic character, ionization energy, atomic size and electronegativity.

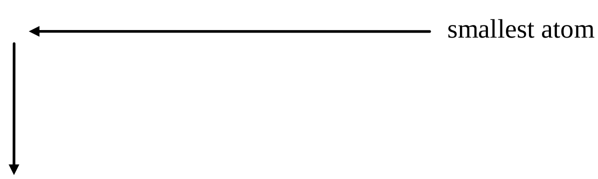
a) **Reactivity in metals and Metallic Character** (the more easily an atom can lose an electron)

This is intended to show that the reactivity in the metals is highest to the left within a period of the periodic table and also highest toward the bottom of a group.

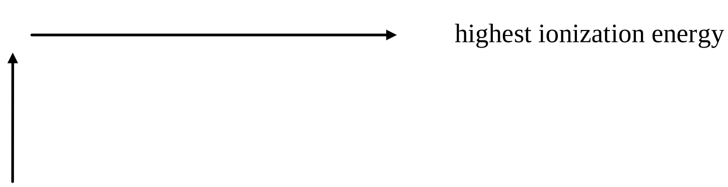
b) **Reactivity in non-metals** (the more easily an atom can gain an electron)



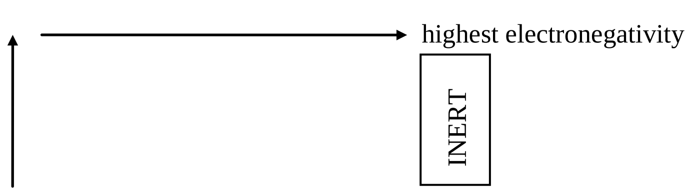
c) **Atomic Radius or Size** (the distance from the nucleus to the valence electrons)



d) **Ionization Energy** (the energy required to remove an electron from an atom)



e) **Electronegativity** (measures the attraction an atom has for the electrons in a bond)



Practice Questions: Trends on the Periodic Table

Trends on the Periodic Table may be explained in terms of the electron arrangements within atoms.
Trends across a period can be explained using the concept of: ______________________________
Trends down a group can be explained using the concept of: ______________________________

1. Consider the elements potassium and rubidium:
a) Are these elements metals or non-metals? ______________
b) Will these elements tend to gain or lose electrons? ______________
c) Which of these elements is more reactive? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.

2. Consider the elements chlorine and bromine:
a) Are these elements metals or non-metals? _______________
b) Will these elements tend to gain or lose electrons? _____________
c) Which of these elements is more reactive? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.

3. Consider the elements sulfur and chlorine:
a) Are these elements metals or non-metals? _______________
b) Will these elements tend to gain or lose electrons? ______________
c) Which of these elements is more reactive? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.

4. Consider the elements potassium and scandium:
a) Are these elements metals or non-metals? _______________
b) Will these elements tend to gain or lose electrons? ______________
c) Which of these elements is more reactive? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.

5. Consider the elements nitrogen and phosphorus:
a) Define first ionization energy.
b) Write the first ionization equation for each of these elements.
c) Which element has the higher first ionization energy? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.

6. Consider the elements carbon and boron:
a) Write the first ionization equation for each of these elements.
b) Which element has the higher first ionization energy? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.

7. Consider the elements gallium and indium. Which of these two atoms has a larger radius? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.

8. Consider the elements zinc and iron. Which of these two atoms has a larger radius? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.

9. Consider the elements chlorine and bromine:
a) Define electronegativity.
b) In general what kind of element, metal or non-metal, has a stronger attraction for new electrons?
c) Which element, chlorine or bromine, has higher electronegativity? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.
d) Electronegativity values are given on the "back" of your periodic table. Check a few of these values to confirm that your answer to the previous question is correct.

10. Consider the elements potassium and scandium. Which element has higher electronegativity? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.