Chapter 14: Chemical Periodicity
Part 1 - Notes: The Periodic Table

Objectives: Identify, define, and explain: periodic law, metal, nonmetal, metalloid, group, family, period, alkali metal, alkaline earth metal, transition metal, inner transition metal, halogen, noble gas, lanthanide series, actinide series, transuranium element, and representative element.
Explain how properties of an element may be inferred based on its position on the periodic table and based on its neighboring elements.
Use electron configurations to classify elements as noble gases, representative, transition, or inner transition.

Text Reference: Section 14.1 – pages 391-396

Elements were most successfully classified by whom?

Originally categorized how?

What is the ***Periodic Law***?

The Modern Periodic Table of Elements

The modern periodic table of elements is arranged according to . . .

Metallic elements are found . . .

Metals always form what type of ions?

Nonmetallic elements are found . . .

What is a metalloid and what can be said about the properties of metalloids?

Vertical columns on the periodic table are called ____________________ or ____________________

Elements in the same family have similar chemical properties. Elements in the same family have ions of the same charge (usually). What is the reason for similar chemical properties and ion charges within families?

Horizontal rows on the periodic table are called ____________________

The number of the period is the number of the highest occupied energy level (*n*).

Group 1 elements are called the ____________________ metals.

Chemical properties: These elements are the most highly reactive metals. They are so reactive they must be kept under a layer of oil to keep them from reacting with air.

Physical properties: They are metals so they are silver-grey in color, they are malleable, ductile, and conduct heat and electricity.

Charge of ions in this group: End of electron configuration for elements in this group:

What is the equation for the reaction of an alkali metal with water?

Group 2 elements are called the ____ metals.

Chemical properties: These elements are very reactive, although not as reactive as group 1 elements.

Charge of ions in this group: End of electron configuration for elements in this group:

Groups 3 - 12 are a set of elements called the ____.

They are given this name because they were thought to be transitional between metals and nonmetals.

What do you know about the charge of ions in this set of elements?

What is the final term of the electron configuration in the elements in this set of elements?

Group 17 elements are called the _________________________.

These are the most reactive of the nonmetallic elements.

Charge of ions in this group: End of electron configuration for elements in this group:

Group 18 elements are called the _________________________.

These are the most unreactive elements on the periodic table. They used to be called ***inert gases***; but this is an incorrect name because inert means nonreactive. The noble gases are highly unreactive but, in some circumstances, they do react.

Charge of ions in this group: End of electron configuration for elements in this group:

The elements in the *Lanthanide Series* are atomic numbers ________ to ________.

Where is this set of elements located?

The lanthanide series is also classified as . . .

The elements in the *Actinide Series* are atomic numbers ________ to ________.

Where is this set of elements located?

The actinide series is also classified as . . .

What is a ***transuranium element?***

What is a ***representative element?***

What groups contain representative elements?

Elements are either representative or transitional.

Question 1: Classify each element as a representative element, a transitional element, or a noble gas.

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

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

c. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

d. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

Question 2: Explain how an element's outer electron configuration is related to its position on the periodic table.

Chapter 14: Chemical Periodicity
Part 1 – Assignment: The Periodic Table

Periodic Table Grid for Identification

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
A																	J
B	C											E	F	G	H	I	K
B	C											E	F	G	H	I	K
B	C	D	D	D	D	D	D	D	D	D	D	E	F	G	H	I	K
B	C	D	D	D	D	D	D	D	D	D	D	E	F	G	H	I	K
B	C	D	D	D	D	D	D	D	D	D	D	E	F	G	H	I	K
B	C	D	D	D	D	D	D	D									
			L	L	L	L	L	L	L	L	L	L	L	L	L	L	
			M	M	M	M	M	M	M	M	M	M	M	M	M	M	

Identify the part of the periodic table to which the statement applies by the letter in the boxes. It may need more than one letter.

1. Noble gases ______________________
2. Halogens ______________________
3. Inner transition metals __________________
4. Contains only metals ____________________
5. Contains only nonmetals __________________
6. s-block ______________________
7. Lanthanide series ______________________
8. Alkaline earth metals ____________________
9. e^{-1} configuration ends in p^4 ________________
10. d-block ______________________
11. e^{-1} configuration ends in s^1 _________________
12. Actinide series ______________________
13. Alkali metals ______________________
14. e^{-1} configuration ends with $5f^x$ _________________
15. p-block ______________________
16. Transition metals ______________________
17. Least reactive ______________________
18. Most reactive metals ______________________
19. Used to be rare earth metals ____________________
20. f-block ______________________
21. Ions of variable charge ______________________
22. Most reactive nonmetals ______________________

Chapter 14: The Periodic Table
Part 2 - Notes: Ion Configurations

Objectives: Write complete and abbreviated electron configurations for anions and cations.
Define isoelectronic and determine and identify isoelectronic species.
Relate the charge, number of valence electrons, and configuration of an ion to its position on the periodic table.
Describe the formation of cations from metals and of anions from nonmetals

Text Reference: Section 15.1 – pages 413-418.

For Representative Elements

Example 1a: Write the electron configuration for **calcium**.

Calcium forms a ____________ ion by _________________________ electrons. Which electrons would be lost by the calcium atom to form its ion? The electrons in its outer most energy level, of course. These would be the __________ electrons.

Example 1b: Write the electron configuration for the **calcium ion**.

Name	*Symbol*	*# Protons*	*# electrons*	*Electron Configuration*
Potassium atom				
Potassium ion				
Argon				
Chlorine atom				
Chloride				

Argon and *potassium ion* are said to be ***isoelectronic***, they have the same electron configuration. *Argon* and *chloride* are also isoelectronic. *Chloride* and *potasium ion* are isoelectronic. NOTE: *Chloride* and *potassium atom* are NOT isoelectronic.

Representative elements often for react to form ions that have electron configurations with full s and p sublevels in the highest level.

For Transition Elements.

Transition elements often form more than one ion. This is because its outer electrons may be lost from both the *s* and the *d* orbitals. Transition elements are ***NOT*** isoelectronic with noble gases. Consider the element iron and its two ions, iron (II) and iron (III).

Name	*Symbol*	*# protons*	*# electrons*	*Configuration*
Iron				
Iron (II)				
Iron (III)				

As a general rule, when ions are formed from transition metals, atoms tend to lose electrons from *sublevels with the higher energy level* (the higher coefficient).

Nickel

Nickel (II)

Nickel (III)

Copper

Copper (I)

Copper (II)

Chapter 14: The Periodic Table
Part 2 – Notes: Ion Configurations

Write the complete electron configurations for the following elements and ions:

1. strontium ion

2. iodide

3. chromium (II)

4. telluride

5. zinc ion

6. mercury (II)

7. manganese (II)

8. cobalt (III)

9. With which noble gas is zinc isoelectronic?

10. With which noble gas is telluride isoelectronic?

For questions 11-21, write the word true or false in the blank provided.

11. ______________________ The strontium ion is isoelectronic with xenon.

12. ______________________ Chloride is isoelectronic with the potassium ion.

13. ______________________ Nitride is isoelectroic with the aluminum ion.

14. ______________________ Nitride is isoelectronic with the magnesium ion.

15. ______________________ Nitride is isoelectronic with oxide.

16. ______________________ Nitride is isoelectronic with oxygen.

17. ______________________ Nitrogen is isoelectronic with oxygen.

18. ______________________ Bromide is isoelectronic with chloride.

19. ______________________ The rubidium ion is isoelectronic with iodide.

20. ______________________ Oxide is isoelectronic with helium.

21. ______________________ The lithium ion is isoelectronic with neon.

Chapter 14: Chemical Periodicity
Part 3 - Notes: Periodic Trends

Objectives: Interpret group trends in atomic radii, ionic radii, ionization energies, and electronegativities.
Interpret period trends in atomic radii, ionic radii, ionization energies, and electronegativities.
Identify, define, and explain: atomic radius, ionic radius, electronegativity, ionization energy, & electron affinity.
Text Reference: Section 14.2 – pages 398-406

Shielding and Effective Nuclear Charge

Shielding occurs because, in a multiple electron atom, each electron acts as a ***shield*** for electrons farther out from the nucleus, reducing the attraction between the nucleus and the distance electrons.

Effective nuclear charge (Z_{eff} or Z*) is a measure of the nuclear attraction for an electron.

- There are two electrons in the 1s orbital next to the nucleus? What do they "feel"? They "feel" the pull of the protons in the lithium nucleus.
- There is a single electron in the second energy level. What does it "feel"? It feels less than the +3 of the protons that the first two electrons "feel". It feels (very) approximately 3 – 2 = +1 charge. It is ***shielded***. The two 1s electrons are "covering" 2 of the 3 protons in the nucleus. So the ***effective nuclear charge*** (***Z_{eff}***) of the third electron of lithium is ~+1.
- This makes sense, since the third electron of lithium is in the second energy level (*n* = 2) and it is further away from the nucleus – and not as tightly held as the electrons in the *1s* orbital.

$Z^* = Z_{actual} - S$

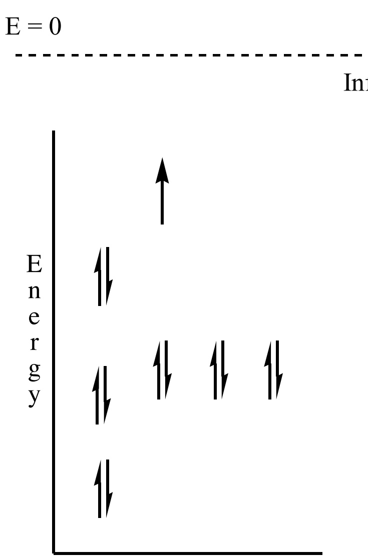
- Z_{actual} is the number of protons and S refers to a shielding constant
- S incorporates the electron-electron repulsive forces and the shielding of lower energy electrons
- It is important to remember that each electron in a multi-electron atom has its own Z* which can be calculated experimentally based on the energy required to remove that specific electron from the atom.

Electrons in the *nd* orbitals are completely shielded by the electrons in the *ns* and *np* orbitals. But, the electrons in the *ns* and *np* orbitals are not completely shielded by the electrons in the *(n-1)s* and *(n-1)p* orbitals.

Ionization Energy

Ionization energy – the energy required to remove the most loosely held electron from an atom or ion in the gaseous state

- IE is the energy required fro the following process: $X_{(g)} \rightarrow X^{+}_{(g)} + e^{-}$
- It is always energy that needs to be put into to accomplish this process – because it takes energy to pull an electron away from a neutral atom. You are separating charges (electrostatic attraction between the positive proton and the negative electron) and the charges do not want to be separated – and of course it takes energy to get this to happen.
- ***First ionization energy:*** energy required to eject the most weakly held electron from an atom resulting in a +1 ion
- Ionization energy INCREASES going from *left to right* across a period in the periodic table.
- Ionization energy INCREASES going *up* a family in the periodic table.



Infinite distance from the nucleus - zero energy

Putting a positive charge next to a negative charge is favorable - so it would be a release of energy to bring them together.

Sio if you pull apart a positive charge and a negative charge, it would **require** energy. Itis considered a **negative energy** value.

So the higher the orbital is, the closer it is to E = 0 and the less energy it would take to get the electron away from the atom.

- The higher the energy of the orbital, the lower the ionization energy.
  - If you could boost the energy of an orbital – it would be easier to pull an electron away from the neutral atom
  - The higher energy of the orbital, the closer it is to being infinity – and easier to pull away
  - For multi-electron systems, you are trying to remove the least strongly held electron – that is the highest energy valence electron.

Ionization energy increases from left to right across a period and increases from bottom to top of a family.
EXCEPTIONS – Boron and Oxygen. WHY???

Be: $1s^2\ 2s^2$ B: $1s^2\ 2s^2\ 2p^1$ B^+: $1s^2\ 2s^2$

- The electron removed from boron is the first one in the *p* orbital and we know (because E $\propto$ the type of orbital) that *p* orbitals are higher in energy than *s* orbitals and it is easier to pull out a *p* electron with higher energy (in boron) than to remove a *2s* electron with lower energy (from Be).

N: $1s^2\ 2s^2\ 2p^3$ O: $1s^2\ 2s^2\ 2p^4$ O^+: $1s^2\ 2s^2\ 2p^3$

Even though oxygen is very electronegative – there is still a drop in the ionization trend. Why?

- When adding the electrons to oxygen's orbitals, the final electron is the fourth *p* electron and is added to the *2p* orbitals. There is a cost to pair the electrons and a cost due to electron-electron repulsion.
- The p^4 electron is slightly higher in energy than the last electron in nitrogen due to those two energy repulsions.
- So to remove such an electron requires slightly less electron because it has slightly higher energy and would be easier to get rid of the final electron in oxygen than the final electron in nitrogen.
- It takes more energy to remove the $2p^3$ electron from nitrogen than it does to remove the $2p^4$ electron from oxygen due to the special stability that arises from half-filled sublevels.
- The same reasoning can be used to explain the deviation between phosphorus and sulfur. The only difference is the principal quantum number of the valence electrons.

In a group, the first ionization energy tends to decrease as atomic number increases.

Core shielding makes removal of valence electrons easier. As the atomic number increases, there is increased shielding of the nucleus due to the core electrons. Since the nucleus is being shielded more, it has less and less pull on the valence electrons.

Successive Ionization Energies

The energy required to remove a second electron from an atom is known as the second ionization energy and is abbreviated IE_2. The energy required to remove a third electron from the orbital is known as the third ionization energy, abbreviated IE_3.

Note that it generally takes a greater amount of energy to remove each additional electron.
Also note that it takes large increases in energy to remove core electrons once all valence electrons have been removed.
It is much easier to remove valence electrons from an atom than it is to remove core electrons.

Atomic Radius

- We know that an orbital has no definite size; it is simply a probability map describing the probable location of an electron. We know that an atom consists of a nucleus and electrons that exit in orbitals, outside the nucleus forming an electron cloud.
- The size of an orbital is not precisely defined, so the size of an atom cannot be precisely defined. However, it is possible to estimate the size of an atom.
- By studying solid crystalline structures of elements and compounds through a technique known as X-ray diffraction, an estimate of the distance between the different nuclei of the crystal can be obtained. Scientists have also been able to estimate the distance between the nuclei of diatomic molecules. By knowing the distance between nuclei of atoms, scientists may estimate the size of the atom.
- ***Atomic Radius*** is half the distance between the nuclei of atoms.

- Going across a period:
  - The Z increases
  - There is the same shielding from the electrons in lower energy levels – but the ones in the outer shell don't shield each other. In fact, electrons in the same energy level don't shield each other very well.
  - The electrons feel a bigger and bigger pull as more protons are added to the nucleus – but since the ones in the outer level are not getting any additional shielding – they are also feeling a bigger pull on the nucleus.
  - Since the electrons are feeling more of a pull – and getting pulled in further – the radius decreases across a period in the periodic table.

- Key: Electrons in higher energy levels do not shield electrons in lower energy levels.
- For Gallium – its valence electrons are sucked down more than you would think
- Going from Ca ($4s^2$) to Gallium ($4p^1$) – you didn't just add 10 electrons - you added 10 electrons that didn't shield very well and are pulled into the nucleus more than you'd expect.
- Atomic radius is roughly proportional to Z* (effective nuclear charge) and to *n*.
- **Question:** Is Z a periodic property? No. It just keeps getting bigger and bigger.
- **Question:** Is Z* (effective nuclear charge) a periodic property? Yes. It is a pattern that repeats throughout the periodic table.

<u>Key Point:</u> *As atomic number increases **in a period**, the atomic radii of the elements in the period tend to **decrease**.*
*As atomic number increases **in a period**, the atomic radii of elements in the group tend to **increase**.*

Elements in a group have similar valence electron configurations. The difference between valence electron configurations of elements in a group is that as the atomic number increases, the *energy level* (n) of the valence electrons increases. The *greater the energy level of the valence electrons, the greater the probability of finding those electrons at greater distances from the nucleus.*

Another factor affecting size of the atom is called **shielding**. *Core electrons shield the valence electrons from the attractive force of the nucleus making it possible for them to move greater distances from the nucleus.*

<u>Atoms to Cations</u>

The electron configuration of calcium is:

The electron configuration of the calcium ion is:

Note that the calcium atom has electron through the fourth energy level. The calcium ion only has electrons through the third energy level. Because of this there is a much greater probability that the electrons in the calcium ion will be closer to the nucleus than the electrons of the calcium atom.

	Number of protons	*Number of electrons*
calcium atom		
calcium ion		

Note that both the calcium atom and the calcium ion have the same number of electrons; however, the calcium ion has two less electrons than the calcium atom. There is **less** electron repulsion in the calcium ion.
Also the 20 protons in the nucleus of the calcium ion can pull the 18 electrons in the ion closer to the nucleus than the 20 protons in the atom can pull the 20 electrons toward the nucleus of that atom.

<u>Cations to Cations</u>

	Symbol	*# electrons*	*# protons*	*ionic radius in nm*	*Electron configuration*
sodium ion					
aluminum ion					

Note that the aluminum ion has a smaller radius than the sodium ion because the greater the number of protons in aluminum ion nucleus may exert a greater force over the lesser electrons in the sodium ion nucleus.

<u>Atoms to Anions</u>

The electron configuration of chlorine is:

The electron configuration of chloride is:

There are more electrons in the outermost orbital of the chloride than in chlorine. Electron repulsion is therefore increased and the electrons move farther away from the nucleus.

	Number of protons	*Numbers of electrons*
chlorine atom		
chlorine ion		

Note that both the chlorine atom and the chlorine ion have the same number of protons. The 17 protons in the chlorine atom can pull the 17 electrons closer than the 17 protons in the ion can pull the 18 electrons.

Anions to Anions

	Symbol	*# electrons*	*# protons*	*ionic radius in nm*	*Electron configuration*
nitride					
fluoride					

Note that there are more protons in the fluoride ion than there are in the nitride ion. The nuclei of both ions are pulling on the same number of electrons. The nucleus with the greatest number of protons is able to pull the electrons closer.

Chapter 14: Chemical Periodicity
Part 3 - Notes: Periodic Trends

Record the answers to the following text questions on this sheet in the appropriate locations. Questions are on pages 409 – 410.

17.

18a. 18b. 18c. 18d.

20a. 20b. 20c.

21.

22a. 22b. 22c.

23.

24a. 24b. 24c.

25.

29.

32a. 32b. 32c.

35.

43.