Chemistry 10: Review of concepts and test

April 20-April 24

Time Allotment: 40 minutes per day

Student Name: ______________________________

Teacher Name: ______________________________

Academic Honesty

I certify that I completed this assignment independently in accordance with the GHNO Academy Honor Code.

Student signature:

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I certify that my student completed this assignment independently in accordance with the GHNO Academy Honor Code.

Parent signature:

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Packet Overview

Date	Objective(s)	Page Number
Monday, April 20	*Review of electron shielding and practice problems.*	3
Tuesday, April 21	*Review of periodic trends*	6
Wednesday, April 22	*Short quiz on electron shielding*	9
Thursday, April 23	*Practice problems re: metals, nonmetals, and metalloids*	13
Friday, April 24	*.Holiday*	

Additional Notes:

Hi students.

This week there's an unusual setup. You'll (again) be walking through parts of the periodic table. This is a review of past material. Then you'll be given two gentle quizzes (one after two days of review, the second after one day of review), so you'll be doing a lot of practice. The second quiz (on metals and atomic character) will be on **Monday, April 27.**

Expect to be tested, and expect to reference your notes and remember thoughts from last week.

Remember, these reflections are a powerful tool for being more familiar with the *urstoff*—the primordial substances that make up our world. You should expect to be developing, slowly, and begin to do more leisurely thinking about elements and their relations.

Please email me at any time, with any questions, and good luck this week.

Mr. Luke

Monday, April 20

Chemistry Unit: Review of concepts and test
Lesson 1: Review: Effective nuclear charge, calculation and theory

Unit Overview

A Question for Lesson 1: What does the equation $Z_{eff} = Z - S$ communicate? What does it tell me about an electron?

Objective: Be able to do this by the end of this lesson.
Using the reference materials and material learned last week, prepare for the quiz on atomic properties. Be comfortable using and understanding the limits of the equation $Z_{eff} = Z - S$.

Agenda to Lesson 1

I. Watch video: calculating Z_{eff} and limitations to the simple model

II. Practice problems:

7.9a
7.10b
7.11
7.12
7.13
7.14

III. Review solutions, check answers, & make corrections in a red pen.

Solutions

7.9

a. When electrons orbit the nucleus, they are being pulled by its charge. The stronger the nucleus' charge, the stronger the pull between the electron and its nucleus. However, some factors (like electron shielding) get in the way and decrease the effect that the nucleus' charge has on the electron. For example, the nucleus may have a charge of 4 (having 4 protons), but because of shielding an electron might experience that as a charge of 3. We call this decreased amount of charge that takes into account outside factors the *effective nuclear charge.*

7.10

b. The 1s electrons. The electrons closer to an atom's nucleus always experience something closer to the full charge of the nucleus when compared to electrons that are farther away. This is because the inner electrons repel the outer electrons and reduce the pulling effect exerted by the nucleus.

7.11

a. Potassium has 19 protons, so it would have a nuclear charge (Z) of 19. Between the 4s orbital layer and the nucleus lie these orbitals: 1s, 2s, 2p, 3s, and 3p. Therefore, The 4s electron is being shielded from the nucleus by 2 + 2 + 6 + 2 + 6 (18) electrons. Using the known equation for estimating Z_{eff},

$$Z_{eff} = Z - S = 19 - 18 = 1$$

Therefore, the effective nuclear charge felt by a 4s electron of potassium is +1 (assuming that electrons are *totally* effective in shielding outer electrons and the valence electrons provide no screening for each other).

If the Z_{eff} for a 4s K electron is actually +3.49, that means that the electrons in those inner orbitals (1s, 2s, 3s, 3p, etc) are not as good as shielding the nucleus from the outer 4s electron as I expected. In other words, they don't totally shield the nucleus from the outer electron.

7.12

a. This is the same problem as before. Sulfur has 16 protons, so it would have a nuclear charge (Z) of 16. Between the 3p orbital layer and the nucleus lie these orbitals: 1s, 2s, 2p, and 3s. Therefore, The 4s electron is being shielded from the nucleus by 2 + 2 + 6 + 2 (12) electrons. Using the known equation for estimating Z_{eff},

$$Z_{eff} = Z - S = 16 - 12 = 4$$

Therefore, the effective nuclear charge felt by a 4s electron of potassium is +4 (assuming that electrons are *totally* effective in shielding outer electrons and the valence electrons provide no screening for each other).

b. Since the Z_{eff} for a 3p Sulfur electron is actually +10.52, it's clear that the inner electrons are not fully shielding the outer electron(s) from the nucleus. However, I also notice that this is a huge difference when compared to the difference between the actual and theoretical Z_{eff} values for the 4s potassium electron (3.59 vs 1, whereas for this electron it's 10.52 vs 4). The solutions manual adds: "All s electrons have a finite probability of being close to the nucleus and inside the core … For this reason, shielding of s electrons is never perfect and calculated values of Z_{eff} reflect this".

7.13

Okay. Both electrons being considered are in the same energy level. Therefore, if they belonged to the same atom, I'd expect the Z_{eff} for an average electron in each case to be the same. However, the atoms are not the same type. Krypton and Argon primarily differ in their atomic number and therefore will, probably, differ somewhat in structure. Argon has an atomic number of 18, and Krypton has an atomic number of 36.

I remember that Z_{eff} has two factors: Z (which is determined only by the number of protons in the nucleus) and S (which is determined by the number of electrons between the electron under consideration and the nucleus). I notice that for Krypton, Z is much larger (36 vs. 18). I notice for S, I'm considering an electron that is in the same energy level in both atoms. In other words, the S is the same for both examples being considered.

Sum it up this way: $Z_{eff\ of\ Ar} = 18 – S$, $Z_{eff\ of\ Kr} = 36 – S$. Clearly Z_{eff} of the Kr electron will be much higher. The Z_{eff} for the n=3 electron in Kr will be higher than the Z_{eff} for the n=3 electron in Argon, sinmply because Krypton has more protons.

I think, therefore, the Krypton n=3 electron will be closer to the nucleus, simply because it experiences a stronger pull to the nucleus than the Argon n=3 electron.

7.14

By the same reasoning above, if I'm considering an electron in the n=3 level (for instance, a 3s electron in a K atom vs. a 3s electron in a Mg atom), **they have the same number of electrons between them and the nucleus** (remember S = # of electrons between the electron being considered and the nucleus). Therefore, the amount of shielding occurring for each electron should be about the same. If that's true, then the charge of the nucleus (the # of protons) is the only thing I need to consider.

Therefore, a n=3 electron will experience the weakest pull, the lowest Z_{eff} in Mg (12 protons in the nucleus).). After that, P (15), K (19), Ti (22), and Rh (45).

Tuesday, April 21

Chemistry Unit: Review of concepts and test
Lesson 2: Review: Periodic trends of atomic properties

Unit Overview

3 questions for Lesson 2: As I move across a row or down a column, what is changing? How do basic increases in protons and electrons in and around an atom influence atomic properties like ionization energy and atomic radius?

Objective: Be able to do this by the end of this lesson.
Using the reference materials and material learned last week, prepare for the quiz on atomic properties (ionization energy, atomic radius, the effect of nuclear charge and electrons in different n=? energy levels and orbitals).

Agenda for Lesson 2: Review of periodic trends

I. Do practice problems:

    7.17
    7.21
    7.22
    7.23
    7.26
    7.37
    7.39
    7.41
    7.42

II. Review solutions, check answers, & make corrections in a red pen.

Solutions

7.17

The phrase "this environment" struck me as odd. What's the environment of this W atom? Well, I know that if a W atom is part of a piece of tungsten (W) metal, then it is a lattice of metal W atoms bound together (see metal bonding, Quarter 2). The big question here is whether the atoms' electrons are overlapping (therefore making their atomic radii smaller) or not (therefore keeping their atomic radii the same as before the bond formed). I remember that in a piece of metal, the atoms are bound to each other in a way that is essentially covalent (check your notes on this). Therefore, the **bonding radius** of W atoms that are bound together is smaller than the radius of a W atom by itself. I can say for certain that the atomic radius of a W atom in this environment is larger than the bonding radius. I can calculate the bonding radius of a W atom as 2.74×10^{-10} m divided by two, so the atomic (non-bonding) radius of a W atom is a little more than that (1.37×10^{-10} m).

7.21

When moving across a row, you increase the number of protons *and* electrons by one each time. Overall, the size of atoms *decreases*. But it's a close battle.

> There are two factors at work here. As the number of protons increases, the pull it is able to exert on the electrons increases (which tends to compress the atom and make it smaller). This overpowers the effect that adding single electrons has. To elaborate: while the increasing number of electrons would tend to make the size of the atom increase, they are filling up in a single energy level (no more than 3 orbitals), and so they don't increase the size of the atom enough to offset the increased pull of the nucleus. Picture an onion. **Moving across a row, it takes a lot of electrons to add a "layer" to the onion.** For example, you've got to add six electrons to fill the p orbital. This goes along with 6 new protons in the nucleus. Then you've got to add ten electrons to fill the d orbital. This goes along with 10 new protons, which offsets the size increase caused by the new layer. **Adding more electrons within a row, apparently, is not enough to overcome the effect of adding protons**. One related point is that the rows correspond to an energy level (e.g. n=3). In the third row, you remain in the n=3 level, so electrons will still be at a distance that corresponds to the n=3 energy level from the nucleus**. Adding more protons to the nucleus draws that whole n=3 layer inwards**.

On the other hand, moving from one row down to the next row (moving down a column), you move from the n=3 to the n=4 level (for example). While you are adding a *lot* of new protons (just as many as the number of electrons you're adding) by moving down one element in a column, the fact that you've added enough electrons to fall into a new layer (the n=4 layer) means you've added **all** the orbitals in the n=3 layer (s, p, and d, in this case) and one orbital in the new layer (an s orbital, for instance). That's at least four new layers to the onion. **While you are getting many more protons which draw all those new electrons tightly to the nucleus, they aren't enough to overcome all these new *layers* of electrons.**[1] The size of atoms increases as you move down a column in the periodic table.

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[1] This is—in large part—a judgment that we also make based on experimental evidence.

7.22
a. See answer to 7.21

b. S, Si, Se, Ge (use the periodic table)

7.23
a. Be, Mg, Ca
b. Br, Ge, Ga
c. Si, Al, Tl

7.26
a. as Z stays constant and the number of electrons increases, the electron-electron repulsions (shielding) increase, the electrons spread apart, and the ions become larger.
b. Going down a column, the increasing average distance of the outer elecrons from the nucleus causes the size of particles (all with the same charge) to increase.

c. The electron configurations of Fe, Fe^{2+}, and Fe^{3+} are $[Ar]4s^2 3d^6$, $[Ar]3d^6$, and $[Ar]3d^5$, respectively. The 4s valence electrons in Fe are ono average father from the nucleus than the 3 electrons, so Fe is larger than Fe^{2+}. Since there are five five possible ways of arranging the d orbital shape, in Fe^{2+} at least one orbital must contain a pair of electrons. Removing one electron to form Fe^{3+} significantly reduces repulsion, increasing the nuclear charge experienced by each of the other d electrons and decreasing the size of the ion.

7.37
a. Ionization energies are always positive quantities because the atom is considered to be the "system". The "system" must gain energy in order to lose an electron; therefore, the IE is always shown as positive.

b. It has one more proton and electron than Oxygen, and it only needs 1 electron to fulfill its octet. Its nucleus is pulling its outer electrons more strongly than oxygen's because it has one more proton than Oxygen. Therefore, it's harder to remove its outermost electron than Oxygen's outermost electron.

c. A second ionization energy is always higher, meaning that it's harder to remove a second electron, for any atom because that electron is likely to be in an orbital layer that's closer to the nucleus (so it would feel a stronger pull), or (if it's in the same orbital layer as the electron that was removed) it is no longer being partially shielded from the nucleus by the first electron (which you've removed at this point).

7.39
a. In general, larger atomic size/radius means lower first ionization energy. Practically, this can be rephrased as "the larger an atom is, the easier it is to remove the outermost electron from it". This is true because for a large atom, the outermost electron is 1) really far from the nucleus in comparison to a smaller atom and 2) is being shielded/blocked from the nucleus by all the electron orbital layers in between it and the nucleus.

b. Helium (smallest atom, least amount of shielding); Francium (largest atom, most amount of shielding). The title for "lowest IE" could go to a larger atom than Francium; however, we don't talk about them much because they're so darn large, they break apart before we can observe them in any way.

7.41
a. Ar b. Be c. Co d. S e. Te

7.42
Only the element with the larger IE is given; see above solutions for explanations.
a. Mo b. N c. Cl d. Rn

Wednesday, April 22

Chemistry Unit: Review of concepts and test
Lesson 2: Quiz

Quiz: Atomic Properties and Periodic Table Trends

Agenda for today:

Take quiz on Google Classroom. If you have opted out of Google Classroom, take the quiz on next page and fill out answer sheet, then submit via preferred method.

Quiz: Atomic properties and per. table trends

Keep a blank periodic table and a piece of blank paper for scratch work on hand.

For written questions in **bold** – **write your answer on the answer sheet** (not on this quiz).

1. Which of these elements has the smallest atomic radius?
    a. Al
    b. Cl
    c. F
    d. K
    e. Ca
2. Which of these elements has the largest atomic radius?
    a. Al
    b. Cl
    c. F
    d. K
    e. Ca
3. Which of these elements has the lowest first ionization energy?
    a. Oxygen
    b. Boron
    c. Fluorine
    d. Beryllium
    e. Carbon
4. Which of these elements has the highest first ionization energy?
    a. Oxygen
    b. Boron
    c. Fluorine
    d. Beryllium
    e. Carbon
5. Which of these elements would it be easiest to remove an electron from?
    a. Oxygen
    b. Boron
    c. Fluorine
    d. Beryllium
    e. Carbon
6. Which of these electrons would it be easiest to ionize by turning into a cation?
    a. Oxygen
    b. Boron
    c. Fluorine
    d. Beryllium
    e. Carbon

7. Select all factors that have a part in determining an atom's size.
    a. the number of electrons an atom has
    b. the number of protons an atom has
    c. the number of energy levels (e.g. the n=3 energy level, the n=4 energy level, etc) its electrons take up
    d. the strength of the nucleus' pull on the atom's electrons
8. Select all factors that have a part in determining the ionization energy of an electron.
    a. the number of electrons an atom has
    b. the number of protons an atom has
    c. the number of energy levels (e.g. the n=3 energy level, the n=4 energy level, etc) its electrons take up
    d. the strength of the nucleus' pull on the atom's electrons (the Z)
    e. the amount of "shielding" that occurs between electrons
    f. the number of electrons between the electron you're looking at and the atom's nucleus
9. **Define the "effective" nuclear charge, in relation to a given electron around an atom's nucleus.**

10. **Write the formula for calculating the effective nuclear charge, using the variables Zeff, Z, S, as well as =, -, or + signs as appropriate.**

11. **Calculate the Zeff for a Fluorine atom's outermost electron.**

12. Would the Zeff for a Fluorine's outermost (2p) electron be higher or lower than for its innermost (1s or 2s) electrons?
    a. higher
    b. lower
13. As you move from one element to the next, left to right across a row of the periodic table, the size of that element's atom should
    a. increase
    b. decrease
14. As you move from one element to the next, down a column of the periodic table, the size of that element's atom should
    a. increase
    b. decrease
15. As you move from one element to the next, left to right across a row of the periodic table, the outermost electron should be _____________ to remove.
    a. easier (lower ionization energy)
    b. harder (higher ionization energy)
16. As you move from one element to the next, down a column of the periodic table, the outermost electron should be _____________ to remove.
    a. easier (lower ionization energy)
    b. harder (higher ionization energy)

17. Which element is the largest?
    a. S
    b. Si
    c. Se
    d. Ge
18. If it takes 5 kJ of energy to remove an electron from a Francium atom, how much energy would it take to remove a second electron from the Francium atom?
    a. 8 kJ of energy (more energy)
    b. 5 kJ of energy (the same amount)
    c. 2 kJ of energy (less)
19. The bonding distance between two tungsten (W) atoms is 2.74 Angstroms. What is the bonding radius?
    a. 2.74 Angstroms
    b. 1.37 Angstroms
    c. 5.44 Angstroms
20. If the bonding radius of sodium is 3.40 Angstroms, you'd expect the atomic (non-bonding) radius of sodium to be
    a. 3.82 (higher)
    b. 3.40 (the same)
    c. 3.29 (lower)
21. **Why do large atoms tend to have very low first ionization energies (why is the first electron very easy to remove for large elements)? Explain in 1-2 sentences and use the terms electron, nuclear charge, and pull in your explanation.**

22. **In a paragraph (3-5 sentences), explain why the size of atoms increases as you move down the periodic table; for instance, from Fluorine to Iodine. Be sure to account for the number of protons, electrons, and how orbitals play a role in this (possibly) contradictory phenomenon.**

Name: __________________ Answer Sheet (submit this only)

1. _____
2. _____
3. _____
4. _____
5. _____
6. _____
7. _____
8. _____
9. ____
   ____
   ____
10. ____
11. __________
12. _____
13. _____
14. _____
15. _____
16. _____
17. _____
18. _____
19. _____
20. _____
21. _____
22. _____

Thursday, April 23

Chemistry Unit: Review of concepts and test
Lesson 3: Review: Atomic properties and the Metals

Unit Overview

Lesson 4 Socratic Question: What makes the metals characteristically themselves? How is this consistent with our knowledge of atomic properties (ionization energy, atomic radius, etc)?

Objective: Be able to do this by the end of this lesson.

Using the periodic table, knowledge of electron configurations, and atomic properties, predict chemical and physical properties about metals in preparation for the quiz tomorrow (or Monday, depending on the holiday schedule).

Agenda for Lesson 4

I. Review concepts and material from last week.

II. See chapter summaries on pages 290 and 291 for reference.

III. Do practice problems.

7.53
7.54
7.55

What is "metallic character"?

7.57
7.58

7.65
7.68
7.71
7.73
7.76

IV. Review solutions, check answers, & make corrections in a red pen.

The quiz for this material will be given on **Monday, April 27.**

Solutions

I've transcribed these solutions directly from the textbook, so these solutions are not as nicely-phrased and targeted to your learning style as last week's solutions. Underline and highlight anything that seems relevant or most important to your answer. There will be some information in two answer that you haven't seen before—simply read and take what you can from them.

7.53
The smaller the first ionization energy of an element, the greater the metallic character of the element. In other words, the more easily an element's outermost electron can be removed/displaced, the more "metallic" it is. This makes sense, because metals' electrons move freely. This is the reason metals "conduct" electricity so easily. Their outermost electrons take part in the flow, and a metal atom loses and gains electrons with great ease when electrons are passed through a piece of metal.

7.54
S is a nonmetal, Si and Ge are metalloids, and Ca is a metal. We expect that electrical conductivity increases as metallic character increases. Since metallic character increases going down a column and to the left in a row, the order must be S < Si < Ge < Ca.

7.55
Metallic character increases moving down a column and moving left-right across a row.
a. Li b. Na c. Sn d. Al

What is "metallic character"?
Metallic character is, in short, the collection of properties that make a metal characteristically a metal. Electrical conductivity, low ionization energy, thermal conductivity, shiny lustre, malleability, ductility, solid at room temperature with one exception (mercury), and high melting points with three exceptions (Hg, Ce, and Ga).

7.57
To solve this one, you'll need to remember that ionic compounds are formed by combining a nonmetal and a metal atom.

With that in mind, your answers are:

Ionic	Molecular (covalent)
MgO	SO_2
Li_2O	P_2O_5
Y_2O_3	N_2O
	XeO_3

7.58
a. Two examples you could use (use your textbook and notes for these)

For metal oxides:

$$Li_2O\ (s) + H_2O\ (l) \rightarrow 2\ LiOH\ (aq)$$

$BaO\ (s) + H_2O\ (l) \rightarrow Ba(OH)_2\ (aq)$

For non-metal oxides:

$P_4O_{10}\ (s) + 6\ H_2O\ (l) \rightarrow 4\ H_3PO_4\ (aq)$

$SO_3\ (g) + H_2O\ (l) \rightarrow H_2SO_4\ (aq)$

b. Metals have lower ionization energies than nonmetals, so they tend to form ionic oxides, while nonmetals form molecular oxides. Ionic compounds, in this case oxides, dissociate into ions when they dissolve into water. The reactive oxide ion ends up as a hydroxide, separated from the metal ion, Molecular oxides do not ionize upon dissolution, so the oxygen remains bound to the nonmetal.

7.65

Na	Mg
$[Ne]3s^1$	$[Ne]3s^2$ Mg has one more electron than Na
+1	+2 Na loses one electron, while Mg must lose two 3s electrons to gain a [Ne] configuration
+496 kJ/mol	+738 kJ/mol The nuclear charge of Mg is greater than that of Na, so it requires more energy to remove a valence electron with the same *n* value from Mg as Na. It also means that the 2s electrons of Mg are held closer to the nucleus, so that the atomic radius is smaller than that of Na.
very reactive	Reacts with steam, but not H_2O (l) Mg is less reactive because it has a filled subshell and it has a higher ionization energy.
1.54 Angstroms	1.30 Angstroms See long explanation above.

7.68

a. Cs is much more reactive than Li toward H_2O because its valence electron is less tightly held (greater *n* value) and Cs is more easily oxidized.

b. The purple flame indicates that the metal is potassium. Reference Figure 7.23 if desired.

c. $K_2O_2\ (s) + H_2O\ (l) \rightarrow H_2O_2\ (aq) + K_2O\ (aq)$
potassium peroxide + water → hydrogen peroxide + potassium oxide

7.71

H: $1s^1$, Li: $[He]2s^1$, F: $[He]2s^22p^5$. Like Li, H has only one valence electron, and its most common oxidation number is +1, which both H and Li adopt after losing the single valence electron. Like F, H needs only one electron to adopt the stable electron configuration of the nearest noble gas. Both H and F can exist in the -1 oxidation state, when they have gained an electron to complete their valence shells.

7.73

F	Cl
[He]$2s^2 2p^5$	[Ne]$3s^2 3p^5$
-1	-1 Although Cl has significantly more electrons, and is in the next energy level (n=3 as opposed to n=2 for Fluorine), it is missing only one electron to gain the nearest noble gas configuration, just like F, so it takes the same ionic charge.
+1681 kJ/mol	+1251 kJ/mol The n=2 valence electrons in F are closer to the nucleus and more tightly held than the n=3 valence electrons in Cl. Therefore, the ionization energy of F is greater, and the atomic radius is smaller.
Reacts exothermically to form HF	Reacts slowly to form HCl In its reaction with H_2O, F is reduced; it gains an electron. Although the electron affinity, a gas phase single atom property of F is less negative than that of Cl, the tendency of F to hold its own electrons (high ionization energy) coupled with a relatively large exothermic electron affinity makes it extremely susceptible to reduction and chemical bond formation. Cl is unreactive to water because it is less susceptible to reduction.
-328 kJ/mol	-349 kJ/mol While F has approximately the same Z_{eff} as Cl, its small atomic radius gives rise to large repulsions when an extra electron is added, so the overall electron affinity of F is smaller (less exothermic) than that of Cl.
0.71 Angstroms	0.99 Angstroms The n=2 valence electrons in F are closer to the nucleus so the atomic radius is smaller than that of Cl.

7.76

Xe has a lower ionization energy than Ne. The valence electrons in Xe are much farther from the nucleus than those of Ne (n=5 vs n=2) and much less tightly held by the nucleus; they are more “willing” to be shared than those in Ne. Also, Xe has empty 5d orbitals that can help to accommodate the bonding pairs of electrons, while Ne has all its valence orbitals filled.