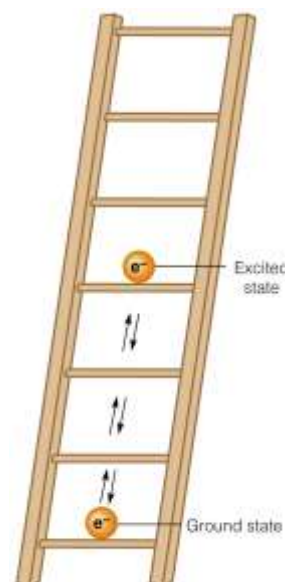


Answers to Homework: The Hydrogen Bright Line Spectrum

Read section 1.4 in your text (Nelson 11, p. 37-40) and answer the following questions:

1. Which two questions could not be answered by Rutherford's model of the atom? (p.37)
 - Rutherford's model could not explain why the protons in the nucleus stayed together. Protons are positively charged so they "should" repel each other and fly apart
 - Rutherford's model also could not explain why the negatively charged electrons would stay in the space around the positively charged nucleus, instead of being attracted to it and spiraling inward
2. What factor determines how much energy an "energy level" possesses?
 - the energy of the energy level depends on how far the electrons in that level are from the nucleus. As the distance from the nucleus increases, the potential energy also increases
3. Summarize the three assumptions made by Bohr in his model of electron arrangement.
 - electrons can travel forever in an energy level without losing energy and slowing down
 - electrons in energy levels further from the nucleus have higher energy (as distance from the nucleus increases, the potential energy of the electrons also increases)
 - electrons can not remain in between orbits. They can absorb energy and move to an energy level further from the nucleus or they can release energy and move to an unfilled orbit closer to the nucleus, but they can not stay in between energy levels.

A good analogy for the Bohr model is the rungs of a ladder. Electrons can only be found on the rungs of the ladder (at specific energy levels). The further they are up the ladder, the higher their potential energy. Electrons can not stay in between the rungs (in between energy levels).



4. Define quantum, ground state, bright line spectrum and continuous spectrum.
 - a) a **quantum** is a unit of energy. It is the amount of energy of one photon of light. An object can have 1,2,3...millions of quanta of energy, but it is impossible to divide a quantum of energy so you can not have a $\frac{1}{2}$ quantum of energy.
 - b) **ground state** is the lowest energy state of an electron in an atom. This is when the electron is as close to the nucleus as possible.
 - c) **bright line spectrum** is the characteristic pattern of bands of light (separated by black, no light regions) given off by atoms of an element when they are excited by the addition of energy. Each colour of light on the bright line spectrum corresponds to an excited electron moving from a specific **higher** energy level to a specific **lower** energy level. For example, an electron dropping from $n = 6$ to $n = 5$ will give off one band of colour on the bright line spectrum, and an electron dropping from $n = 6$ to $n = 4$ will give off a separate band of colour on the spectrum.
 - d) **continuous spectrum** is like a rainbow- it contains all of the different colours of light. The colours run together and overlap. There are no black light regions between the colours of light.

Answers to 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?
 - it is the arrangement of the electrons in an atom that determines the atom’s chemical and physical properties (how the atom will behave)
 - b) Which characteristic of Bohr’s model of the atom is now known to be incorrect?
 - Bohr believed that atoms move in precise, defined orbits around the nucleus and he believed that all of the electrons in an energy level were moving in exactly the same orbit, exactly the same distance from the nucleus (ie. all electrons in an orbit had the same potential energy)
 - c) Give two reasons why electrons do not travel in precise, defined orbits.
 - electrons have wave-like properties. Electrons behave in some ways like a particle and in other ways like a wave, so they do not follow a perfect, predictable orbit.
 - electrons repel each other. Electrons are constantly being attracted toward the nucleus and trying to get as close to it as possible, but at the same time they are being repelled by other electrons. So, whenever one electron moves, it repels the electrons close to it and this changes their motion.
 - d) State Heisenberg’s Uncertainty Principle. What does it mean?
 - Heisenberg’s Uncertainty Principle states that we can never know BOTH where an electron is (its energy, or distance from the nucleus) AND its motion (trajectory or momentum) at the same time. If we design an experiment to find out where an electron is (for example, have it hit something to determine where it is), then we change where it is going. If we try to find out where an electron is going (its motion or trajectory), we will change where it is. So instead of a perfect orbit in which we know BOTH where an electron’s position and motion (where it is and where it is going), we can only describe the regions in an atom where we are likely to find a certain electron. This 3-D region is called an orbital.
 - e) Define “orbital”. How is an orbital different than an orbit?
 - an **orbit** is a precise, defined, measurable path that an object is travelling on; we know exactly where an object is and where it is going. For example, Haley’s comet orbits the sun. At any time, we know BOTH how far Haley’s comet is from the sun (its potential energy with respect to the sun) and where the comet is going (its trajectory or motion).
 - An **orbital** is a three-dimensional area in the space around the nucleus of an atom where an electron is likely to be found greater than 90% of the time. An orbital is a not precise or defined. Sometimes, the electron is not in this space. So, orbitals do not tell us for sure where an electron will be or where it will be later, but it gives us a good idea where to start looking for it.
 - f) How is the quantum mechanical model of the atom the same as Bohr’s model? How is it different?
 - the quantum mechanical model and Bohr’s planetary model of the atom are the same in that both models have electrons moving in discrete quantum levels around the nucleus of the atom (that is, both models include principal quantum levels, “n”).
 - the quantum mechanical model and Bohr’s planetary model are different because Bohr believed that electrons move in precise, defined, predictable paths that he called orbits, similar to the precise, defined paths that the planets follow as they orbit the sun. Bohr believed that we always know exactly where an electron is, and we can predict where the electron will be at other times. The quantum mechanical model recognizes that electron motion is much more complex than a

simple orbit, because the electrons have wave-like properties and electrons repel each other. Therefore, we never know precisely where an electron is going to be, but we can predict a region where it is likely to be found most of the time. The three-dimensional region where an electron is likely to be found more than 90% of the time is called an “orbital”.

- also, Bohr believed that all of the electrons in a certain principal quantum level were moving on exactly the same orbit so they were all exactly the same distance from the nucleus and had the same potential energy. In the quantum mechanical model, each principal quantum level is subdivided into different sub-levels (types of orbitals). The electrons in each sub-level have different motion and potential energy. That is, electrons in the s, p, d and f orbitals have different patterns of motion and are different distances from the nucleus (different potential energies) from one another.

g) How many electrons can an orbital hold?

- an orbital can hold 2 electrons, with opposite spin

h) What is “n”?

- “n” is the principal quantum number, it is always a whole number. It tells us how far an electron is from the nucleus of the atom. As the value of “n” increases, the electrons in that quantum level are further from the nucleus so they have higher potential energy.

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 2 n = 2 8 n = 3 18 n = 4 32 n = 5 50 In general? $2n^2$

b) What is the **total number of orbitals** in the quantum level when: (remember, an orbital can hold only 2 electrons)

n = 1 1 n = 2 4 n = 3 9 n = 4 16 n = 5 25 In general? n^2

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

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

d) How many:

- s-orbitals are there in the second principal energy level: 1
- p-orbitals are there in the second principal energy level: 3
- d-orbitals are there when n = 3: 5 (if there are “d” orbitals, there are always 5 of them)
- s-orbitals are there when n = 4: 1 (there is one “s” orbital in each quantum level)
- f-orbitals are there when n = 5: 7 (if there are “f” orbitals, there are always 7 of them)

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

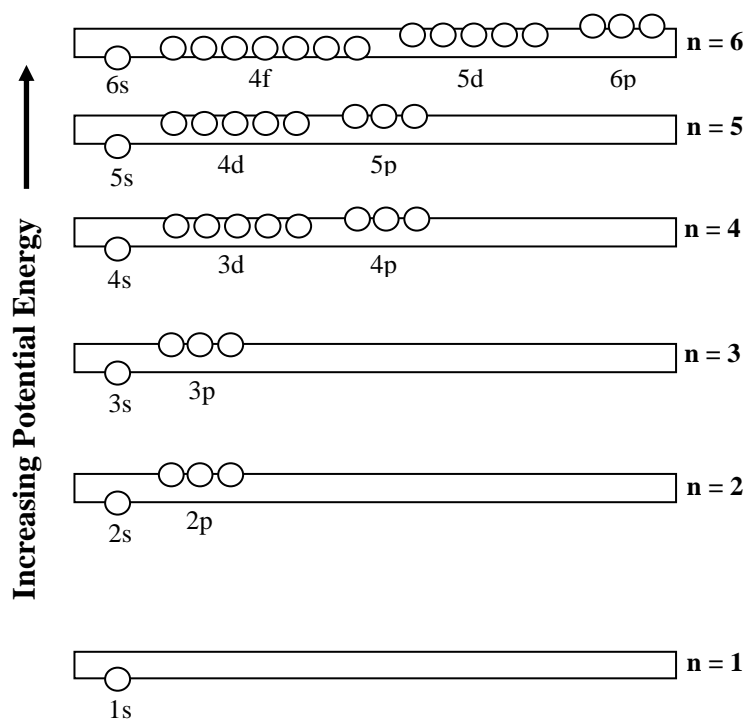
3p 6 4d 10 5f 14 3d 10
3s 2 1s 2 2p 6 4f 14

3. Sample Test Questions:

- a) What is the maximum number of electrons in quantum level 4? 32 When $n=3$? 18 $n=5$: 50
- b) How many electrons can be designated as 3d 10, 4s 2, 5f 14, 2p 6
- c) How many **types** of orbitals are there in quantum level 3? 3 When $n=2$? 2 $n=5$ 5
- d) How many orbitals are there in quantum level 2? 4 When $n=5$? 25 $n=3$ 9
- e) How many electrons can be held in quantum level 5? 50 When $n=1$? 2 $n=4$ 32
- f) How many orbitals are there in quantum level 1? 1 When $n=4$? 16 $n=3$ 9
- g) How many **types** of orbitals are there in quantum level 5? 5 When $n=2$? 2, $n=6$: 6
- h) How many electrons can be designated 2p 6, 4p 6, 5s 2, 6f 14
- i) Circle the orbitals which **do not** exist: $\textcircled{3f}$ 2p 5s $\textcircled{2d}$ 4f $\textcircled{1d}$ 5p 3d $\textcircled{1p}$ 3s 4d 4s

Ways to Remember the Order of Filling for Electron Orbitals

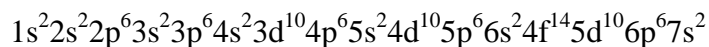
1. Energy Level Orbital Diagrams can be used to remember the order of filling for electron orbitals.



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

7s ²			
6s ²	6p ⁶	6d ¹⁰	
5s ²	5p ⁶	5d ¹⁰	5f ¹⁴
4s ²	4p ⁶	4d ¹⁰	4f ¹⁴
3s ²	3p ⁶	3d ¹⁰	
2s ²	2p ⁶		
1s ²			

Follow the arrows to get the order of filling:



(this is the same order that you get using the energy level diagram, above.)

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

1s ¹	1s ²											"p" block					
2s ¹	2s ²											2p ¹	2p ²	2p ³	2p ⁴	2p ⁵	2p ⁶
3s ¹	3s ²											3p ¹	3p ²	3p ³	3p ⁴	3p ⁵	3p ⁶
4s ¹	4s ²	"d" block										4p ¹	4p ²	4p ³	4p ⁴	4p ⁵	4p ⁶
5s ¹	5s ²	4d ¹	4d ²	4d ³	4d ⁴	4d ⁵	4d ⁶	4d ⁷	4d ⁸	4d ⁹	4d ¹⁰	5p ¹	5p ²	5p ³	5p ⁴	5p ⁵	5p ⁶
6s ¹	6s ²	5d ¹	5d ²	5d ³	5d ⁴	5d ⁵	5d ⁶	5d ⁷	5d ⁸	5d ⁹	5d ¹⁰	6p ¹	6p ²	6p ³	6p ⁴	6p ⁵	6p ⁶
7s ¹	7s ²	6d ¹	6d ²	6d ³	6d ⁴	6d ⁵	6d ⁶	6d ⁷	6d ⁸	6d ⁹	6d ¹⁰	7p ¹	7p ²	7p ³	7p ⁴	7p ⁵	7p ⁶
"s" block																	
		4f ¹	4f ²	4f ³	4f ⁴	4f ⁵	4f ⁶	4f ⁷	4f ⁸	4f ⁹	4f ¹⁰	4f ¹¹	4f ¹²	4f ¹³	4f ¹⁴		
		5f ¹	5f ²	5f ³	5f ⁴	5f ⁵	5f ⁶	5f ⁷	5f ⁸	5f ⁹	5f ¹⁰	5f ¹¹	5f ¹²	5f ¹³	5f ¹⁴		
		"f" block															

Answers to Homework on Electron Configurations

₁₀ Ne	$1s^2 2s^2 2p^6$
₁₁ Na	$1s^2 2s^2 2p^6 3s^1$
₁₂ Mg	$1s^2 2s^2 2p^6 3s^2$
₁₃ Al	$1s^2 2s^2 2p^6 3s^2 3p^1$
₁₄ Si	$1s^2 2s^2 2p^6 3s^2 3p^2$
₁₅ P	$1s^2 2s^2 2p^6 3s^2 3p^3$
₁₆ S	$1s^2 2s^2 2p^6 3s^2 3p^4$
₁₇ Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$
₁₈ Ar	$1s^2 2s^2 2p^6 3s^2 3p^6$
₁₉ K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
₂₀ Ca	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
₂₅ Mn	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
₃₄ Se	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$
₄₀ Zr	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$
₅₀ Sn	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$
₅₅ Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
₆₀ Nd	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^4$
₇₈ Pt	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^8$
₈₅ At	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^5$
₉₂ U	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^4$

(Nd could also be written: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 5d^1 4f^4$)

Similarly, uranium could also be written:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 6d^1 5f^3$ because the $6d^1$ electron goes in before the $5f^1$)

Answers to 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

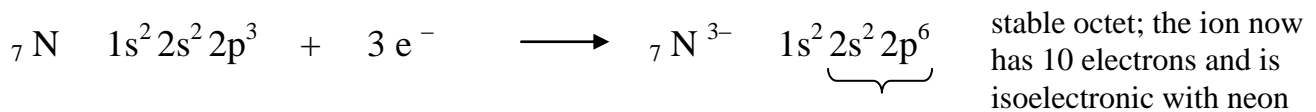
Groups →	1	2											13	14	15	16	17	18
Periods (rows)	n = 1																	
	n = 2																	
	n = 3			3	4	5	6	7	8	9	10	11	12					
	n = 4																	
	n = 5																	
	n = 6																	
	n = 7																	

2. "Read" the following electron configurations to determine the identities of the following elements:

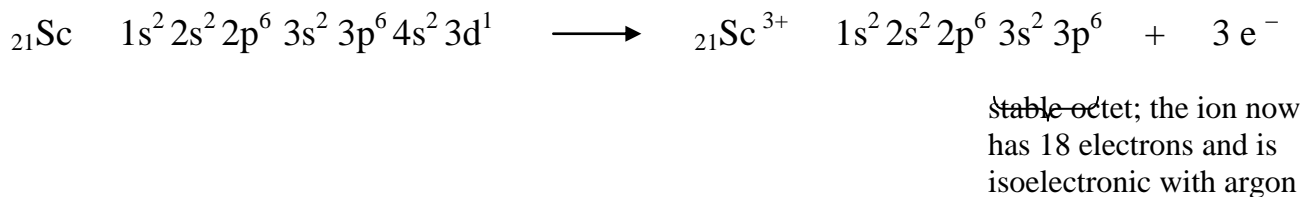
- | | |
|---|------------|
| a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$ | strontium |
| b) $1s^2 2s^2 2p^6 3s^2 3p^3$ | phosphorus |
| c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$ | radon |
| d) $1s^2 2s^2 2p^2$ | carbon |
| e) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$ | iodine |

3. Write ionization reactions for the following elements:

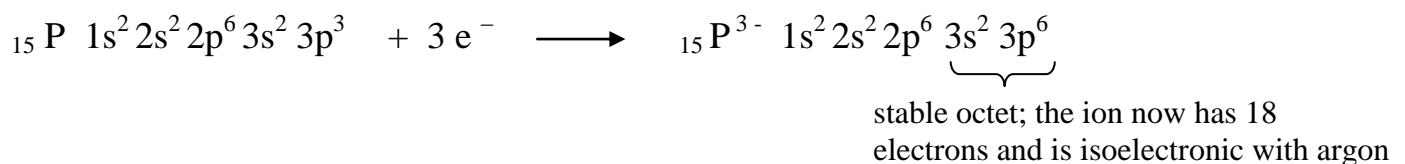
- a) N is a non-metal, so it will gain electrons. It needs 3 more electrons to achieve a stable octet, so it will gain three electrons. Show gained electrons on the **left-hand** side of the arrow.



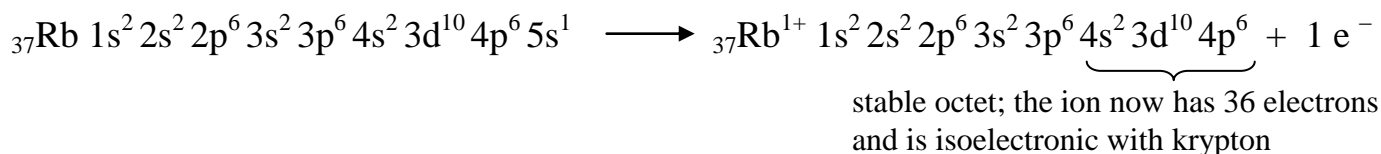
- b) Sc is a metal, so it will lose electrons. It needs to lose 3 electrons to achieve a stable octet. Show lost electrons on the **right-hand** side of the arrow.



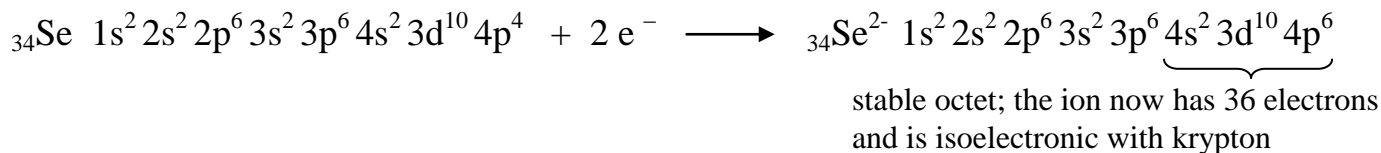
- c) P is a non-metal, so it will gain electrons. It needs 3 more electrons to achieve a stable octet, so it will gain three electrons. Show gained electrons on the **left-hand** side of the arrow.



- d) Rb is a metal, so it will lose electrons. It needs to lose 1 electron to achieve a stable octet. Show the lost electron on the **right-hand** side of the arrow.



- e) Se is a non-metal, so it will gain electrons. It needs 2 more electrons to achieve a stable octet, so it will gain two electrons. Show gained electrons on the **left-hand** side of the arrow.



4. Explain what is meant by the term “isoelectronic”.

The prefix “iso” means “the same”, so isoelectronic means that two atoms or ions have the same number of electrons or the same electron configuration.

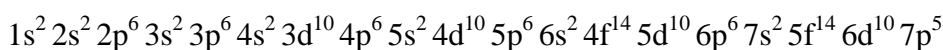
5. What are four ions, with their charges, that are isoelectronic with each of the following:

- a) neon: N^{3-} O^{2-} F^{1-} Na^{1+} Mg^{2+} Al^{3+} c) krypton: As^{3-} Se^{2-} Br^{1-} Rb^{1+} Sr^{2+} Y^{3+}
 b) argon: P^{3-} S^{2-} Cl^{1-} K^{1+} Ca^{2+} Sc^{3+} d) a S^{2-} ion: P^{3-} Cl^{1-} K^{1+} Ca^{2+} Sc^{3+}

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

No, calcium has not turned into argon. Remember, it is the atomic number (number of protons) that determines what type of element an atom is. The calcium ion has the same number of electrons as an argon atom. This just means that they both have a stable octet with 18 electrons.

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.



This element belongs to the Halogen family or Group 17 (VIIA). It will gain one electron to complete a stable octet, so it will form ions with a charge of 1-.

Answers to 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: **net nuclear attraction (NNA or Z_{eff})**

Trends down a group can be explained using the concept of: **shielding effect**

1. Consider the elements potassium and rubidium:

- Are these elements metals or non-metals? **metals**
- Will these elements tend to gain or lose electrons? **lose**
- Which of these elements is more reactive? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.
 - rubidium is more reactive
 - potassium and rubidium both need to lose one electron to achieve a stable octet
 - both elements are in the same group, so they both have the same net nuclear attraction. There is only a net 1+ attraction holding the valence electrons to the nucleus
 - because rubidium has a larger shielding effect, its valence shell is further from the nucleus. Its nucleus will exert a weaker attraction on its valence electron, so it will lose this electron more easily, making rubidium more reactive

2. Consider the elements chlorine and bromine:

- Are these elements metals or non-metals? **non-metals**
- Will these elements tend to gain or lose electrons? **gain**
- Which of these elements is more reactive? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.
 - chlorine is more reactive
 - chlorine and bromine both need to gain one electron to achieve a stable octet
 - both elements are in the same group, so they both have the same net nuclear attraction. There is a net 7+ attraction holding the valence electrons to the nucleus and attracting a new electron
 - because chlorine has a smaller shielding effect, its valence shell is closer to the nucleus. Its nucleus will exert a stronger attraction for a new electron, making chlorine more reactive

3. Consider the elements sulfur and chlorine:

- Are these elements metals or non-metals? **non-metals**
- Will these elements tend to gain or lose electrons? **gain**
- Which of these elements is more reactive? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.
 - chlorine is more reactive
 - sulfur and chlorine both need to gain electron(s) to achieve a stable octet
 - both elements are in the same period, so they both have the same shielding effect. The valence electrons are essentially the same distance from the nucleus so this can not explain the difference in reactivity
 - sulfur has a net nuclear attraction of 6+, so it is attracting new electrons with a 6+ "pull". Chlorine has a net nuclear attraction of 7+, so it has a stronger attraction for a new electron. Because of its stronger NNA, chlorine will have a stronger attraction for a new electron so it is more reactive

4. Consider the elements potassium and scandium:
- Are these elements metals or non-metals? **metals**
 - Will these elements tend to gain or lose electrons? **lose**
 - Which of these elements is more reactive? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.
 - potassium is more reactive
 - potassium and scandium both need to lose electron(s) to achieve a stable octet
 - both elements are in the same period, so they both have the same shielding effect. The valence electrons are essentially the same distance from the nucleus so this can not explain the difference in reactivity
 - potassium has a net nuclear attraction of 1+, so it is holding its valence electron with only a 1+ “pull”. Scandium has a net nuclear attraction of 3+, so it is holding its valence electrons with a 3+ “pull”. Because of its weaker NNA, potassium’s valence electron is only weakly attracted to the nucleus, so it will be lost more easily than scandium’s valence electrons, making potassium more reactive

5. Consider the elements nitrogen and phosphorus:

- Define first ionization energy.

First ionization energy (IE1) is the amount of energy required to remove one electron from a neutral atom of an element.

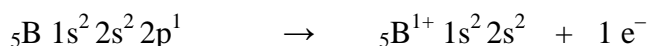
- Write the first ionization equation for each of these elements.



- Which element has the higher first ionization energy? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.
 - nitrogen has the higher first ionization energy
 - both nitrogen and phosphorus are in Group VA (Group 15), so they have the same net nuclear attraction (5+), so this can not be used to explain any differences in ionization energy
 - because nitrogen has a smaller shielding effect, its valence shell is closer to the nucleus. Its nucleus will exert a stronger attraction on its valence electrons, making it more difficult to remove a valence electron from nitrogen than it would be from phosphorus

6. Consider the elements carbon and boron:

- Write the first ionization equation for each of these elements.



- Which element has the higher first ionization energy? Use the concepts of shielding effect and net nuclear attraction (Z_{eff}) to explain why.
 - carbon has the higher first ionization energy
 - carbon and boron are both in the second period, so they have the same shielding effect. Shielding effect can not be used to explain any differences in ionization energy
 - boron is in Group IIIA. It has a net nuclear attraction of 3+, so it is “holding” its valence electrons with a 3+ “pull”. Carbon is in Group IVA. It has a net nuclear attraction of 4+, so it is “holding” its valence electrons with a 4+ “pull”. Because of its stronger net nuclear attraction, it will be harder to remove a valence electron from carbon, so it has a higher first ionization energy.

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.
- indium has a larger atomic radius
 - indium and gallium are both in Group 13 (IIIA) on the Periodic Table. This means that they have the same net nuclear attraction, so net nuclear attraction can not explain the difference in radius
 - indium is in the fifth period and gallium is in the fourth period. Indium has more full energy levels (electron shells) so its valence electrons are further from the nucleus (it has a greater shielding effect) so indium is larger than gallium
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.
- iron and zinc are both in the same period, so they have the same shielding effect. This will not affect the size of the atoms
 - zinc has a stronger net nuclear attraction than iron. This means that the nucleus of the zinc atom has a stronger “pull” on its valence electrons. The valence electrons in the zinc atom are pulled in closer to the nucleus, which makes the zinc atom smaller than the iron atom

9. Consider the elements chlorine and bromine:

a) Define electronegativity.

Electronegativity is a measure of an atom's ability to attract the electrons in a bond. The electronegativity of atoms is measured relative to fluorine, the atom which has the highest electronegativity.

b) In general what kind of element, metal or non-metal, has a stronger attraction for new electrons?

Non-metals are found further to the right on the Periodic Table, which means that they have a stronger net nuclear attraction than metals. Because they have a stronger net nuclear attraction (Z_{eff}), non-metals have a stronger attraction for extra electrons (except the Noble gases, which have essentially no 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.

- chlorine has a stronger electronegativity than bromine
- both elements are in Group VIIA (Group 17) on the Periodic Table. They have the same net nuclear attraction, so this can not be used to explain differences in electronegativity.
- chlorine is in Period 3, while bromine is in Period 4. This means that chlorine has a smaller shielding effect and its valence shell is closer to the nucleus. Because the valence shell is closer to the nucleus, the nucleus will exert a stronger attraction for a new electron, giving chlorine a higher electronegativity.

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.

If you look on the back of the Periodic Table, in the top right hand corner for each element you will see the electronegativity of that element. The EN of chlorine is 3.16, while the EN of bromine is only 2.96. These numbers do agree with our answer to question 8c.

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.

- scandium has higher electronegativity
- scandium and potassium are both found in the fourth period of the Periodic Table, so they have the same shielding effect. Shielding effect can not be used to explain the difference in electronegativity
- potassium is in Group IA, so it has a net nuclear attraction of $1+$. It has only a very weak attraction for its own valence electron or for a new electron. Scandium is in Group IIIB, so it has a net nuclear attraction of $3+$. It has a stronger attraction for its own valence electrons and for the electrons in a bond, which means that it has a stronger electronegativity.