Periodicity – *It is All About the Trends con't*
ELECTRONEGATIVITY This can be thought of as the ability of atoms to attract a shared electron pair in a covalent bond. Think tug of war! Again, use effective nuclear charge and Coulomb's Law to guide your thinking! As you move from Hydrogen to Lithium to Sodium to Potassium (i.e. down Group I or any group) the atoms electronegativity is **DECREASING**. Why? • Think back to **Coulomb's Law**: $U_E \propto \frac{q_1 q_2}{r}$ As the distance (r) increases, (when you add more shells), the electrostatic potential energy of the electrons decreases. So, the ability of that atom's nucleus to attract electrons from another atom in a covalent bond DECREASES as well since the electrons being attracted are further from the positive nucleus and are further shielded by the inner electrons from the nucleus. As you move from Lithium ***ACROSS PERIOD 2*** to Fluroine the *electronegativity* **INCREASES**. Why? • As explained earlier… It is all about the nucleus! All of the elements in Period 2 have their valence electrons in the same energy shell. However, Li has 3 protons, Be 4, B 5, and so on up to F, which has 9. • The higher the nuclear charge (the more effective the nuclear charge Z_{eff}), the more tightly held the electron, the more attraction the nucleus will have on electrons in another atom in a covalent bond.

Trendy Thoughts
First recognize what the question is asking, i.e. what property is in question and what elements are being compared.
To best answer the question use the following 3 steps…
1. Locate both elements on the periodic table and note the energy shell and subshell of their valence electrons.
2. Compare their values! ▪ If same *energy shell*: argue with Z_{eff} ▪ If different *energy shell*: argue based on distance from the nucleus – for Pete's sake it is all about Coulomb's Law!
3. ENERGY! Comment of whether the electrons are more attracted or less attracted; or whether it takes more energy or less energy to remove the electron, etc… NEVER FORGET the ENERGY relationship! Again – Coulomb's Law!
Be careful when comparing isoelectronic species – those that have the same electron configurations, such as Ca^{2+}, K^+, and Cl^-. If they all have the same number of electrons in the same energy shell then the property differences can be explained using effective nuclear charge! Determine the number of protons and answer accordingly.

NMSI SUPER PROBLEM

Answer the following questions about hydrogen atoms, H, hydrogen molecules, H_2, and hydrogen compounds.

(a) An atom of hydrogen emits a discrete wavelength of electromagnetic radiation at 486 nm as the electron transitions from a higher energy level back to the second energy level $n = 2$. Calculate the energy change, in Joules, associated with this transition.

$\Delta E = h\nu = \frac{hc}{\lambda} = \frac{(6.63\times10^{-34}\text{ J s})(3.0\times10^{17}\ \frac{\text{nm}}{\text{s}})}{486\text{ nm}} = 4.09\times10^{-19}\text{ J}$	**1 point** for correct substitution **1 point** for the correct number of joules (or kJ); unit must be shown to receive credit

(b) A second discrete emission line in an atom of hydrogen has a wavelength of 434 nm as the electron transitions from a higher energy level back to the second energy level ($n = 2$).

(i) Would the change in energy associated with the 434 nm spectral line be greater than, less than, or equal to, that of the 486 nm spectral line? Justify you answer.

The energy change for the 434 nm emission will be greater than that of the 486 nm. $\Delta E = h\nu = \frac{hc}{\lambda}$ Since ΔE and λ are inversely related a decrease in wavelength cooresponds to an increase in energy.	**1 point** for correct answer with justification

(ii) Would the energy level (n) of the electronic transition responsible for the 434 nm spectral line be at a higher energy level or a lower energy level than that of the 486 nm spectral line? Justify you answer.

If the energy change is greater for the 434 nm spectral line then the electron must be at a higher initial energy level since both return to the same energy level, $n = 2$.	**1 point** for correct answer with justification

In a separate experiment a molecule of hydrogen, H_2, absorbs a photon of electromagnetic radiation with a wavelength of 300 nm. The energy required to break the bond in H_2 gas is 432 kJ mol^{-1}.

(c) Calculate the frequency of the photons with a 300 nm wavelength.

$\nu = \frac{c}{\lambda} = \frac{(3.0\times10^{17}\text{ nm s}^{-1})}{300\text{ nm}} = 1.00\times10^{15}\text{ s}^{-1}$	**1 point** for correct calculation

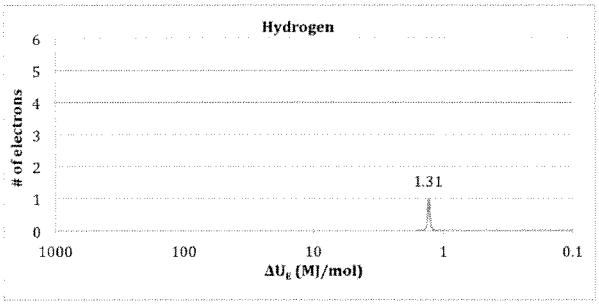
(d) Calculate the number of joules required to break the bond in a single molecule of H_2 gas.

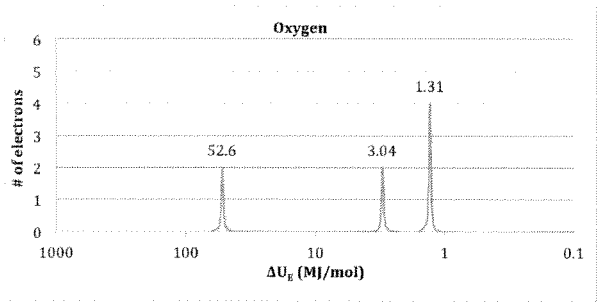
$432\text{ kJ mol}^{-1}\times\frac{1000\text{ J}}{\text{kJ}}\times\frac{1\text{ mol}}{6.02\times10^{23}\text{ molecules}} = 7.18\times10^{-19}\text{ J for each molecule}$	**1 point** for the correct answer

(e) Does the photon have enough energy to break the bond in a molecule of H_2 gas? Mathematically justify your answer.

$E = h\nu = (6.63\times10^{-34})(1.00\times10^{15}) = 6.63\times10^{-19}$ J 6.63×10^{-19} J $< 7.18\times10^{-19}$ J Therefore the photon does not have enough energy to break the bond.	**1 point** for calculating the energy of the photon **1 point** for comparing the energy of the photon to the energy of the single molecule of H_2 and stating the correct answer

The photoelectron spectrographs for both hydrogen and oxygen are shown below.





(f) Using evidence from the PES data, explain why two atoms of hydrogen and one atom of oxygen are required to form water, H_2O.

The PES data shows that H has one electron in its valence shell and oxygen has 4 – so it can form 2 single bonds, one with each of the 2 atoms of hydrogen.	**1 point** for the correct answer referencing the data provided

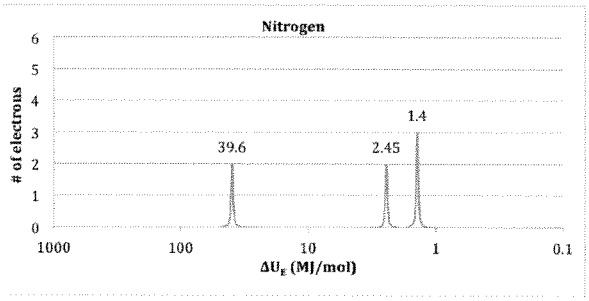
(g) Look at the photoelectron spectrum for oxygen. Circle the peak that contains the first electron that would be removed from an oxygen atom.

The peak labeled 1.31 (the farthest to the right) should be circled as one of those 4 electrons would be the first removed.	**1 point** for correct peak

(h) Give the value for the first ionization energy for an atom of oxygen.

1.31 MJ/mol	**1 point** for the correct first ionization energy

(i) Below is the photoelectron spectrum for nitrogen. Explain why oxygen has a lower first ionization energy than does nitrogen.



Although oxygen has a more effective nuclear charge than nitrogen, oxygen has 2 electrons paired in its *p* orbital. Nitrogen has all three of its *p* electrons unpaired. The paring of oxygen's electrons causes an increase in electron – electron repulsion which decreases the electrostatic potential energy of those 2 paired electrons thus it takes less energy to remove the first one of them.	**1 point** for the correct explanation

(j) Circle the peak in the photoelectron spectrum for nitrogen that represents the electrons in the 2*s* subshell.

The peak labeled 2.45 (the farthest to the right) should be circled.	**1 point** for the correct peak

(k) When hydrogen and nitrogen react they form ammonia. Write the balanced equation for this reaction.

$3\ H_2 + N_2 \rightarrow 2\ NH_3$	**1 point** for the balanced equation

(l) Atoms of phosphorus and nitrogen are in the same group, or family, on the periodic table.

(i) Predict the formula that results when hydrogen atoms form a compound with phosphorus atoms.

PH_3	**1 point** for the correct formula

(ii) Atoms of phosphorus are larger than atoms of nitrogen. Explain

Atoms of phosphorus have valance electrons is the 3rd shell whereas the valence electrons in nitrogen are in the 2nd shell. The more shells the less attracted the electrons are to the nucleus the further away they are from the nucleus.	**1 point** for the correct answer with justification

Modified AP® CHEMISTRY
SCORING GUIDELINES

(4 points)

Answer the following questions related to sulfur and one of its compounds.

(a) Consider the two chemical species S and S^{2-}.

(i) Write the electron configuration (e.g., $1s^2\,2s^2\ldots$) of each species.

S: $1s^2\,2s^2\,2p^6\,3s^2\,3p^4$ S^{2-}: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$ Note: Replacement of $1s^2\,2s^2\,2p^6$ by [Ne] is acceptable.	One point is earned for the correct configuration for S. One point is earned for the correct configuration for S^{2-}.

(ii) Explain why the radius of the S^{2-} ion is larger than the radius of the S atom.

The nuclear charge is the same for both species, but the eight valence electrons in the sulfide ion experience a greater amount of electron-electron repulsion than do the six valence electrons in the neutral sulfur atom. This extra repulsion in the sulfide ion increases the average distance between the valence electrons, so the electron cloud around the sulfide ion has the greater radius.	One point is earned for a correct explanation.

(b) The S^{2-} ion is isoelectronic with the Ar atom. From which species, S^{2-} or Ar, is it easier to remove an electron? Explain.

It requires less energy to remove an electron from a sulfide ion than from an argon atom. A valence electron in the sulfide ion is less attracted to the nucleus (charge +16) than is a valence electron in the argon atom (charge +18).	One point is earned for the correct answer with a correct explanation.

Modified AP® CHEMISTRY SCORING GUIDELINES

(3 Points)

Using principles of atomic and molecular structure and the information in the table below, answer the following questions about atomic fluorine, oxygen, and xenon, as well as some of their compounds.

Atom	First Ionization Energy ($kJ\ mol^{-1}$)
F	1,681.0
O	1,313.9
Xe	?

(a) Write the equation for the ionization of atomic fluorine that requires 1,681.0 $kJ\ mol^{-1}$.

$F(g) \rightarrow F^{+}(g) + e^{-}$	One point is earned for the correct equation. (Phase designations are not required.)

(b) Account for the fact that the first ionization energy of atomic fluorine is greater than that of atomic oxygen. (You must discuss both atoms in your response.)

In both cases the electron removed is from the same energy level ($2p$), but fluorine has a greater effective nuclear charge due to one more proton in its nucleus (the electrons are held more tightly and thus take more energy to remove).	One point is earned for recognizing that the effective nuclear charge of F is greater than that of O.

(c) Predict whether the first ionization energy of atomic xenon is greater than, less than, or equal to the first ionization energy of atomic fluorine. Justify your prediction.

The first ionization energy of Xe should be less than the first ionization energy of F. To ionize the F atom, an electron is removed from a $2p$ orbital. To ionize the Xe atom, an electron must be removed from a $5p$ orbital. The $5p$ is a higher energy level and is farther from the nucleus than $2p$, hence it takes less energy to remove an electron from Xe.	One point is earned for a prediction based on size and/or energy level.

(6 points)

(a) The isotopes have the same number (34) of protons, ***1 pt.***
but a different number of neutrons. ***1 pt.***

- No comment about the number of electrons is necessary

(b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$
or ***1 pt.***
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^4$

- No point is earned for [Ar] $4s^2\ 3d^{10}\ 4p^4$, because the question specifically asks for a complete electron configuration.

Since there are three different $4p$ orbitals, there must be two unpaired electrons. ***1 pt.***

Notes: The second part should have some explanation of Hund's rule, and may include a diagram. The second point can still be earned even if the first point is not IF the electron configuration is incorrect, but the answer for the second part is consistent with the electron configuration given in the first part.

(c) (i) The ionized electrons in both Se and Br are in the same energy level, but Br has more protons than Se, so the attraction to the nucleus is greater. ***1 pt.***

Note: There should be two arguments in an acceptable answer -- the electrons removed are from the same ($4p$) orbital ***and*** Br has more protons (a greater nuclear charge) than Se.

(ii) The electron removed from a Te atom is in a $5p$ orbital, while the electron removed from an Se atom is in a $4p$ orbital. The $5p$ orbital is at a higher energy than the $4p$ orbital, thus the removal of an electron in a $5p$ orbital requires less energy. ***1 pt.***