During Class Invention

Question: Can we use our simple shell model of the atom to make some predictions?

1. Describe the nature of the interaction between protons and electrons in an atom? Consider using some or all of the following terms in your description: attraction, repulsion, neutral, positive, negative, charge, distance, nucleus, force, energy, Coulomb’s Law.

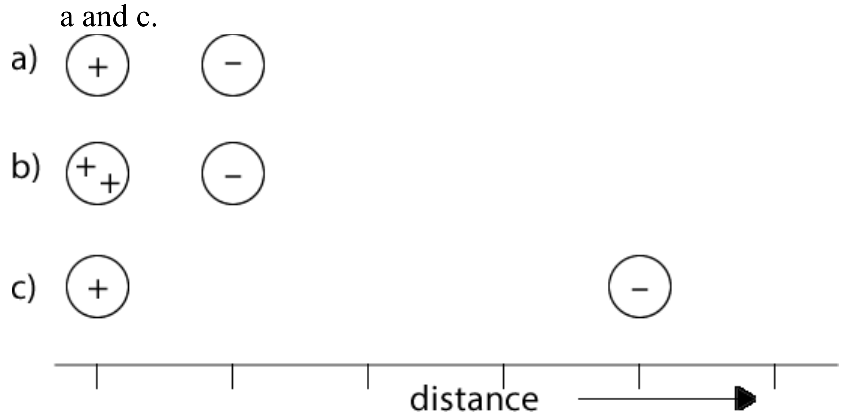
The attraction experienced by the negatively charged electrons to the positively charged nucleus is Coulombic in nature and is described by the equation;

$$\mathbf{E\ \alpha\ \frac{q_1 \cdot q_2}{d}}$$

The equation indicates the energy required to separate the electron (q_1) from the nucleus is directly proportional to the charge on the nucleus (q_2), and inversely proportional to the distance (‘d’) that separates the electron from the nucleus

NOTE: It is interesting when given the option students will choose to use those words that they feel they understand. So some of the words in the list may not appear and under those circumstances we might assume the student does not really have a good understanding of the term.

2. Compare the relative energy necessary to separate positive and negative electrical charges in the following situations? Compare a and b, then compare a and c.

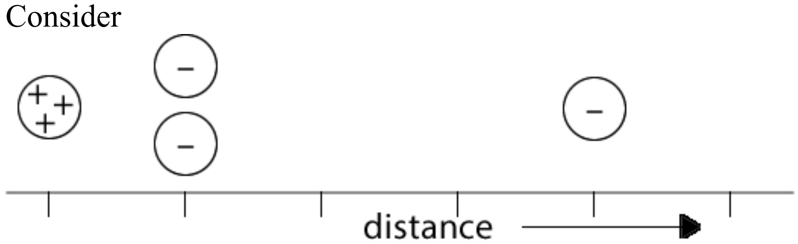


NOTE: The negative charge is the same distance from the positive charge in both a) and b), so the distance term in Coulomb’s Law (E α $\frac{q_1 \cdot q_2}{d}$) is the same for both. However, the positive charge is greater in a) compared to b). The energy required to separate the negative charge from the positive charge in b) is greater

than the energy required to separate the negative charge from the positive charge in a).

NOTE: The negative charge is at a different distance from the positive charge in both a) and c), so the distance term in Coulomb's Law (E α $\frac{q_1 \cdot q_2}{d}$) is important. The positive charge in a) and c) are the same. The energy required to separate the negative charge from the positive charge in a) is greater than the energy required to separate the negative charge from the positive charge in c).

3. Consider



a) how many electrons do you see in the picture? **3 electrons** How many protons? **3 protons**

b) which of these electrons is the easiest (requires the least amount of energy) to remove (ionize)?

The electron furthest to the right (furthest from the nucleus) will be the easiest to remove.

c) Explain your response in b.

According to Coulomb's Law (E α $\frac{q_1 \cdot q_2}{d}$) the energy of attraction (required for separation) is inversely proportional to the distance the electron is from the nucleus.

d) compare the energy from 3b with the energy in 2a and then in 2c.

When comparing 3b to 2a we can think the following way; the electron that is furthest from the nucleus in 3b only experiences an attraction of a net positive charge of +1 because the two electrons in the inner core shield two of the three positive charges in the nucleus from the outer most electron. So the electron in 3b is further from a net +1 positive charge compared to the electron in 3a, which is also experiencing a net +1 charge on its nucleus. Therefore the energy required to remove the electron in 3b is smaller compared to the energy required to remove the electron in 2a.

When comparing 3b to 2c the electron that requires the least amount of energy to remove is approximately the same distance from the nucleus. Also notice that while the nuclear charge in 3b greater, there are also two electrons that 'shield' some of that nuclear charge from the electron the furthest from the nucleus. The net

result is the energy required to remove the electron that is the furthest from the nucleus in 2c and 3b is about the same.

Useful Questions:

What are the similarities and difference between the images in 2a and 2b?
Which (2a or 2b) requires more energy to remove the electron?
Why is less energy required to remove the electron in 2b compared to 2a?
Why is more energy required to remove the electron in 2a compared to 2b?

What are the similarities and difference between the images in 2a and 2c?
Which (2a or 2c) requires more energy to remove the electron?
Why is less energy required to remove the electron in 2c compared to 2a?
Why is more energy required to remove the electron in 2a compared to 2c?

It is VERY important that students always mention both charge and distance in their explanations. This is critical, so even though the distance is the same in 2a and 2b they must point that out in their argument. Just as they must mention that the nuclear charge is the same in 2a and 2c in their argument. Students must know that stating the obvious is most often important in identifying data to support a claim.

The first ionization energy is defined as the minimum energy that must be added to a neutral atom, in the gas phase, to remove an electron from an atom. This definition can be represented in the following chemical equation;

$$\text{energy} + A(g) \rightarrow A^{+}(g) + 1e^{-}$$

4. In the ionization equation above, which is at lower energy? A(g) or $A^{+}(g)$ and $1e^{-}$? Which is at higher energy? A(g) or $A^{+}(g)$ and $1e^{-}$? Explain.

A(g) is at lower energy since according to the equation energy must be added to A(g) to remove the electron. $A^{+}(g)$ and $1e^{-}$ is higher in energy.

5. Explain why energy is required (an endothermic process) to remove the electron in a neutral atom.

Since opposite charges are attractive in nature energy is required to separate oppositely charge particles.

6. The value of the first ionization energy for hydrogen is 1312 kJ mol^{-1}. In the graph below use a short horizontal line to indicate the energy of H(g) (reactant) and a short horizontal line to indicate the energy of $H^{+}(g) + 1e^{-}$ (product). (NOTE: Be sure to consider your responses to Q4 and Q5 above.)

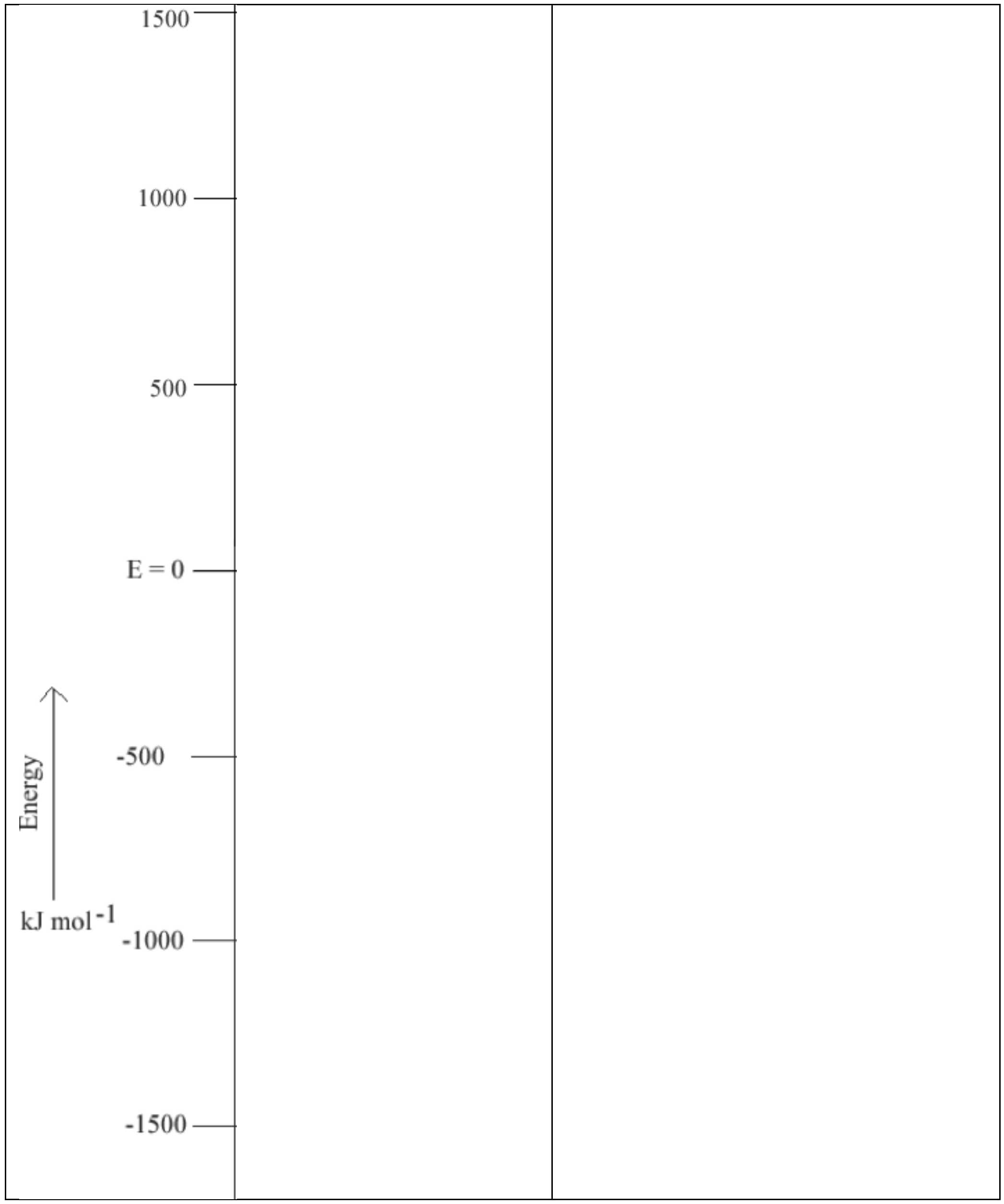
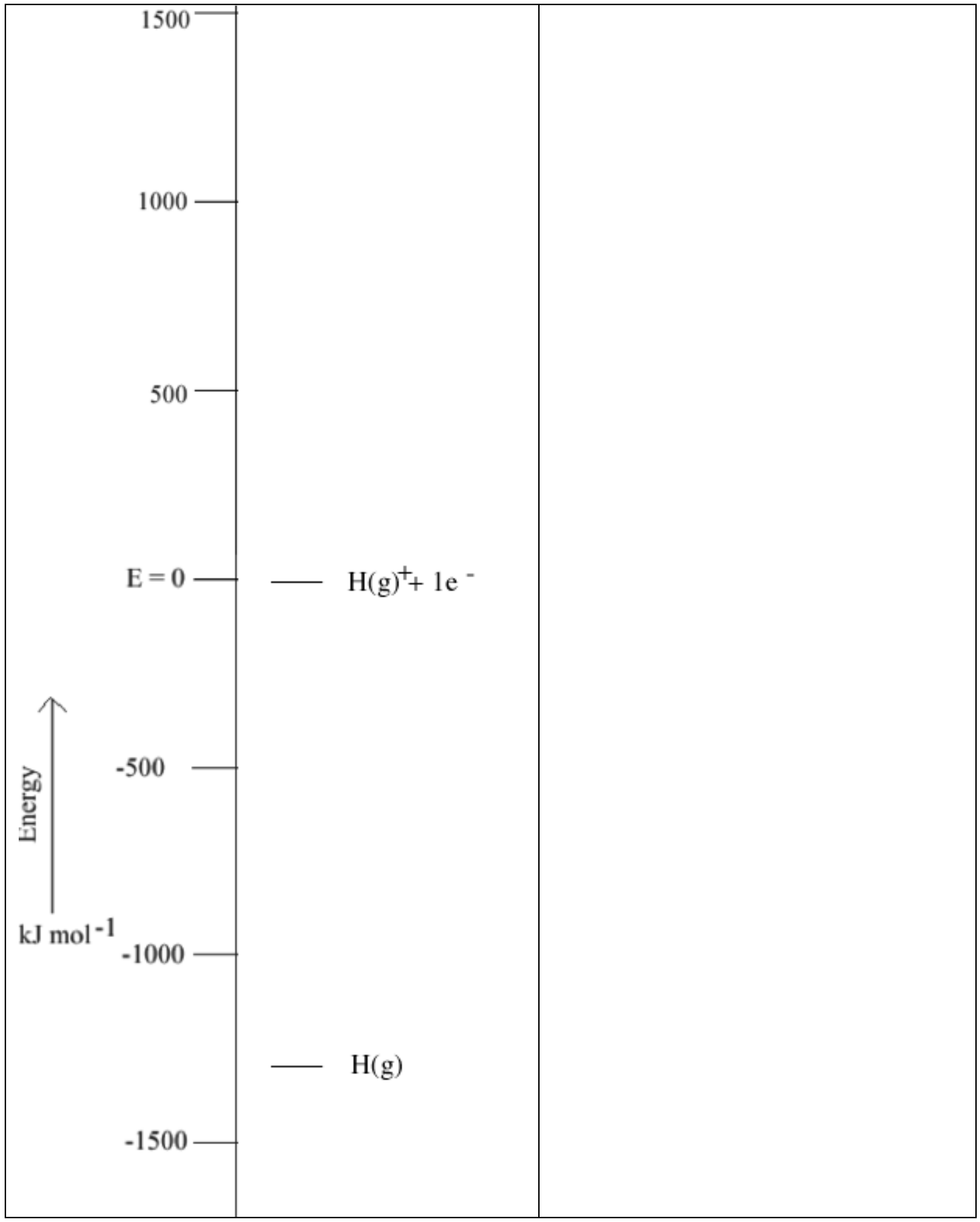


Figure I.

This is a critical point in the discussion. Students know the line for H(g) is lower in energy compared to the line for H^{+}(g) + 1e^{-} . Also the line for the energy of H(g) and the line for energy for H^{+}(g) + 1e^{-} are 1312 kJ mol^{-1} apart, but they do not know where on the y-axis to draw each line. We only know the energy that separates the two lines. To arrive at where the lines are drawn I explain to the students that when the electron is completely removed from the nucleus we will define that as a system of zero energy. When the electron is an infinite distance from the nucleus that that situation is zero energy. Now as the electron moves closer and closer to the nucleus at some point there is an attraction between the electron and the nucleus. The attraction results in a lower energy system, so the line for the H(g) is at -1312 kJ mol^{-1}.



7. What does the difference in energy in the lines in your diagram above represent?

The energy difference represents the ionization energy for the hydrogen atom in the gas phase.

8. The units of energy in Figure I are kJ mol^{-1}. Convert the energy the electron has in a hydrogen atom to J atom^{-1}. In Figure II draw the two horizontal lines

for the energy of the electron in a single hydrogen atom (H(g)) and the energy of the hydrogen ion and one electron ($H^+(g) + 1e^-$).

$$\mathbf{-1312\,\frac{kJ}{mol}\left(\frac{1000\ J}{1\ kJ}\right)\left(\frac{1\ mol}{6.02\times10^{23}\ atom}\right) = -2.18\times10^{-18}\,\frac{J}{atom}}$$

In the energy diagram below draw two horizontal lines that can be used to represent the first ionization energy (J $atom^{-1}$) for hydrogen.

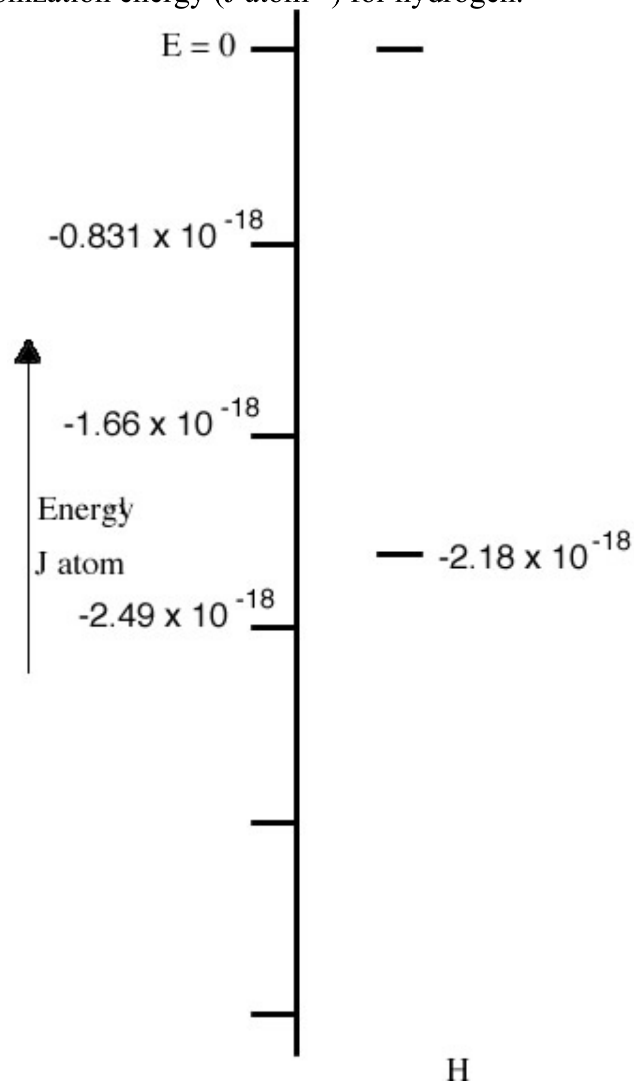


Figure II.

9. How much energy (J $atom^{-1}$) is require to remove the electron from a hydrogen atom (H(g))?

To get from $-2.18 \times 10^{-18}\,\frac{J}{atom}$ to $0\,\frac{J}{atom}$ requires $2.18 \times 10^{-18}\,\frac{J}{atom}$

10. Calculate the wavelength in nanometers of a photon of light capable of ionizing a electron from a hydrogen atom (H(g)).

$$\Delta \mathbf{E} = \frac{\mathbf{h \cdot c}}{\boldsymbol{\lambda}}$$

$$\boldsymbol{\lambda} = \frac{\mathbf{h \cdot c}}{\Delta \mathbf{E}} = \frac{\mathbf{6.626 \times 10^{-34}\ J \cdot s \cdot 3.00 \times 10^{8}\ \frac{m}{s}}}{\mathbf{2.18 \times 10^{-18}\ \frac{J}{atom}}} = \mathbf{9.12 \times 10^{-8}\ m}$$

$$\mathbf{9.12 \times 10^{-8}\ m\ \left(\frac{1 \times 10^{9}\ nm}{1\ m}\right) = 91.2\ nm}$$

This is the wavelength of a photon of light capable of ionizing the electron in a hydrogen atom in the gas phase.

11. Figure III shows an emission spectrum produced by hydrogen atoms in the gas phase.

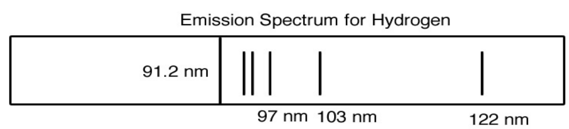


Figure III

12. What region of the electromagnetic spectrum are photons of light with the wavelengths shown in Figure III?

Wavelengths between 90 and 120 nm are in the ultraviolet region of the electromagnetic spectrum.

13. What is the significance of the line at 91.2 nm?

As seen in Q10 above 91.2 nm is the wavelength of a photon that is capable of ionizing a hydrogen atom. So when a hydrogen atom in the gas phase absorbs a photon the electron is ionized, removed.

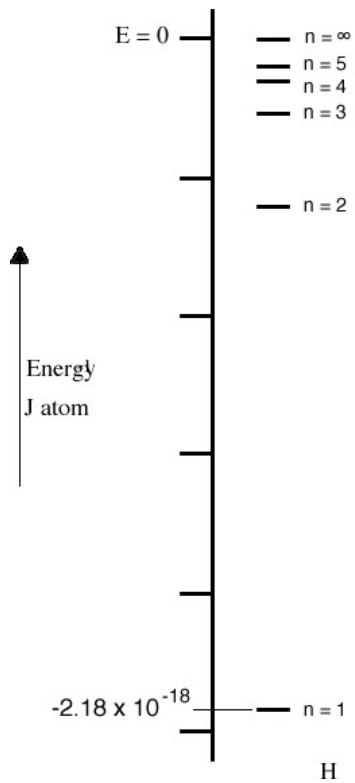
14. The two lines between the line at 91.2 nm and 97 nm have wavelengths of 93.8 nm and 95 nm. Compared to the line at 91.2 nm, the energy of the photons emitted by the hydrogen atom, as depicted in Figure III, are greater than, less than or equal to the energy of the photon at 91.2 nm?

$$\mathbf{93.8\ nm \left(\frac{1\ m}{1 \times 10^{9}\ nm}\right) = 9.38 \times 10^{-8}\ m\ =}$$

$$\Delta E = \frac{h \cdot c}{\lambda} = \frac{6.626 \times 10^{-34}\ J \cdot s \cdot 3.00 \times 10^{8}\ \frac{m}{s}}{9.38 \times 10^{-8}\ m} = 2.12 \times 10^{-18}\ \frac{J}{atom}$$

The energy of this photon is less than the energy required to ionize the electron from the hydrogen atom. The other wavelengths 95 nm, 97 nm, 103 nm and 122 nm are all longer wavelengths and therefore will also have lower energies compared to 2.18 x $10^{-18}\ \frac{J}{atom}$.

15. In a previous activity we used the first ionization energy data as evidence to support a shell model for atoms. In the space below draw an energy level diagram that depicts the energy levels for the first five shells.



16. We have indicated that at E = 0 in the energy level diagram the shell value (n) is infinity. We also know that energy required to remove the electron (that is in the n = 1 shell) from a hydrogen atom has a value of 2.18 x 10^{-18} J. What if a smaller amount of energy is absorbed by a hydrogen atom?

So if adding 2.18 x 10^{-18} J will excite the electron from n =1 shell to the n = ∞ shell, if a smaller amount of energy is absorbed than possibly the electron is excited from the n = 1 shell into one of the shells of higher energy, like the n = 2, n = 3, n = 4, …...

17. What might the electron do if the energy absorbed by the atom is smaller than the amount of energy required to remove the electron?

So if adding 2.18 x 10^{-18} J will excite the electron from n =1 shell to the n = ∞ shell, if a smaller amount of energy is absorbed than possibly the electron is excited from the n = 1 shell into one of the shells of higher energy, like the n = 2, n = 3, n = 4, …...

18. How much energy does a photon of light that has a wavelength of 122 nm have? 103 nm? 97 nm? 95 nm? 93.8 nm?

Energy for a photon of 93.8 nm

$$\Delta E = \frac{h \cdot c}{\lambda} = \frac{6.626 \times 10^{-34}\ J \cdot s \cdot 3.00 \times 10^{8}\ \frac{m}{s}}{9.38 \times 10^{-8}\ m} = 2.12 \times 10^{-18}\ \frac{J}{atom}$$

Energy for a photon of 95 nm is $2.09 \times 10^{-18}\ \frac{J}{atom}$

Energy for a photon of 97.4 nm is $2.05 \times 10^{-18}\ \frac{J}{atom}$

Energy for a photon of 102.5 nm is $1.94 \times 10^{-18}\ \frac{J}{atom}$

Energy for a photon of 122 nm is $1.63 \times 10^{-18}\ \frac{J}{atom}$

19. Can you match up the energy of each of these photons with possible transitions between different shells? Use the energy level diagram below to label these transitions.

See Figure IV below where the transition of the electron moving from the n = 1 shell to the n = 2 shell is assigned the lowest energy, is 1.63 x $10^{-18}\ \frac{J}{atom}$**. The transition from n = 1 to n = 3 is assigned an energy of 1.94 x** $10^{-18}\ \frac{J}{atom}$**. The transition from n = 1 to n = 4 is assigned an energy of 2.05 x** $10^{-18}\ \frac{J}{atom}$**. The transition from n = 1 to n = 5 is assigned an energy of 2.09 x** $10^{-18}\ \frac{J}{atom}$**.**

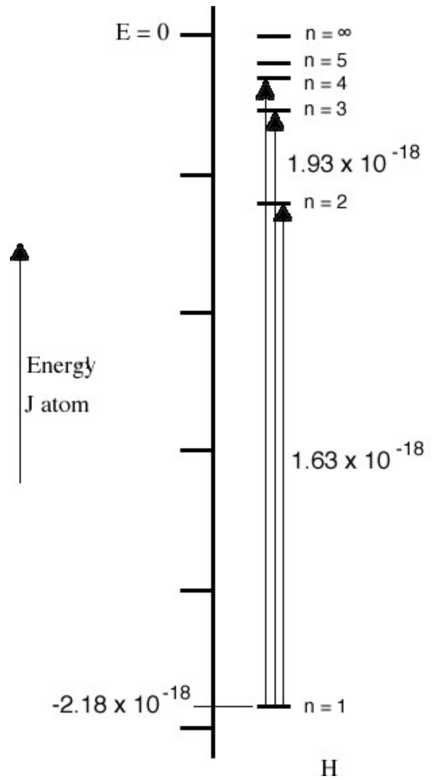


Figure IV

20. What was the assigned energy of the photon you believe is required to excite the electron from n = 1 level to the n = 2 level?

The transition of the electron moving from the n = 1 shell to the n = 2 shell is assigned the lowest energy, is 1.63 x 10^{-18} $\frac{\text{J}}{\text{atom}}$.

21. Knowing the energy of the electron in the n = 1 shell and the energy of the photon required to excite the electron from the n = 1 shell to the n = 2 shell, what is the energy of the n = 2 shell?

The energy of the n = 1 shell is -2.18 x 10^{-18} J, and the energy required to excite the electron to the n = 2 shell is 1.63 x 10^{-18} J, so the energy of the n = 2 shell must be;

-2.18 x 10^{-18} J + 1.63 x 10^{-18} J = -5.5 x 10^{-19} J

22. What is the energy of the n = 3 shell? The n = 4 shell? The n = 5 shell?

The energy of the n = 3 level must be,

-2.18 x 10^{-18} J + 1.94 x 10^{-18} J = -2.4 x 10^{-19} J

The energy of the n = 4 level must be,

$$-2.18 \times 10^{-18}\ J + 2.04 \times 10^{-18}\ J = -1.4 \times 10^{-19}\ J$$

The energy of the n = 5 level must be,

$$-2.18 \times 10^{-18}\ J + 2.09 \times 10^{-18}\ J = -9 \times 10^{-20}\ J$$

23. Determine the energy of the n = 2 through n = 5 levels in terms of value of the n = 1 level and n. Try different combinations of -2.18×10^{-18} J and n to obtain the energy of the other levels. What is the relationship that you come up with?

So the best way to determine how n and -2.18×10^{-18} J is to setup an algebraic expression, like

$$-5.5 \times 10^{-19}\ J = -2.18 \times 10^{-18}\ J \cdot x$$

so x is

$$x = \frac{-5.5 \times 10^{-19}\ J}{-2.18 \times 10^{-18}\ J} = 0.25$$

A value for x of 0.25 is equivalent to multiplying -2.18×10^{-18} J by $\frac{1}{4}$ which would be the same as $\frac{1}{2^2}$. So an equation that combines -2.18×10^{-18} J and n would be

$$E = -2.18 \times 10^{-18}\ J\ \frac{1}{n^2}.$$

We can check this equation with each of the other levels. So for the n = 3 level,

$$E_3 = -2.18 \times 10^{-18}\ J\ \frac{1}{n^2} = -2.18 \times 10^{-18}\ J\ \frac{1}{3^2} = -2.18 \times 10^{-18}\ J \cdot \frac{1}{9}$$

$$E_3 = -2.42 \times 10^{-19}\ J$$

This energy value for the n = 3 shell agrees with the value we determined previously.

24. In the space below draw an energy level diagram that depicts the energy levels for the first five shells and include the energy for each of these levels.

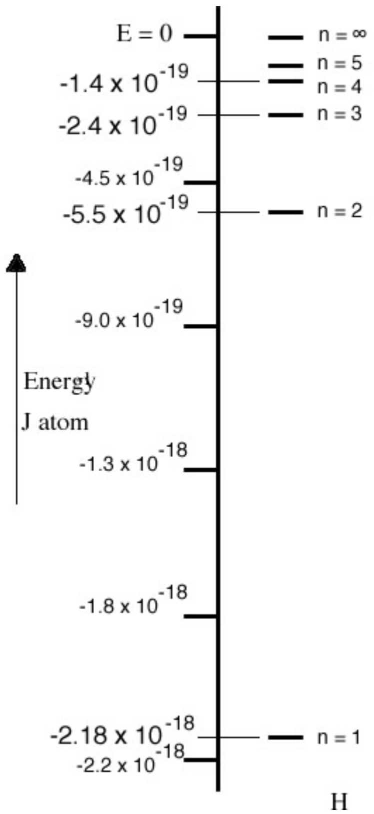


Figure V

25. The energy level diagram drawn above represents, for the hydrogen atom, the energy for the first five shells. These energies were determine based on experimental evidence from emission spectra in the ultraviolet region of the electromagnetic spectrum. The lines in the emission spectrum can be understood as being produced when an excited electron in a higher energy level drops to the n = 1 level, and at the same time releasing a photon with an energy that is the difference between the two levels.

Predict some additional lines that could be produced as a result of an excited electron undergoing different electronic transitions. Determine the region of the electromagnetic spectrum the lines would appear. Explore the internet to determine if your predicted emission lines have actually been experimentally verified and when they were verified and by whom.

In this section students should come up with the idea that other transitions can be produced. Lines in the ultraviolet region are due to excited electrons falling from n = 6, 5, 4, 3 or 2 levels to the n = 1 level. There are two additional sets of lines observed in the visible region and the infrared region. The emission lines in the visible region are produced when excited electrons fall from the n = 6, 5, 4 or 3 levels to the n = 2 level. The emission lines in the infrared region are produced when excited electrons fall from the n = 6, 5, or 4 levels to the n = 3 level.

Students should be able to calculate the ΔE for each of these transitions and then convert the energies of the photons emitted to wavelengths and then google the wavelengths and verify their predictions. Successful verification is the basis of strengthening a model.