Answer on Question #66834 – Chemistry | Inorganic Chemistry

Calculate ionization energy of hydrogen atom in kJ*mol^{-1}.

Solution:

Use the Rydberg expression:

The wavelength λ of the emission line in the hydrogen spectrum is given by:

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

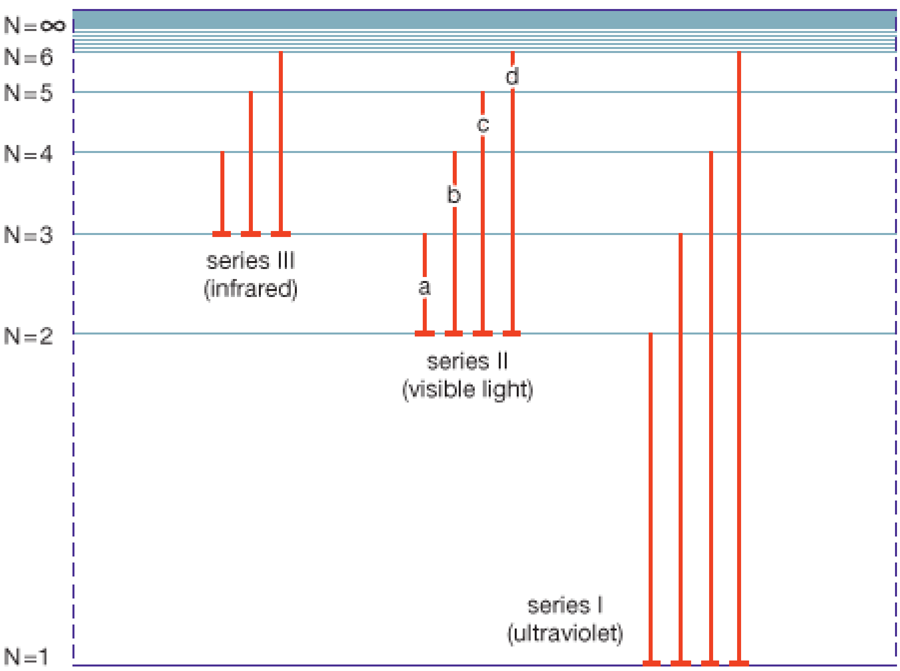
R is the Rydberg Constant and has the value 1.097×107m^{-1}

n_1 is the principle quantum number of the lower energy level

n_2 is the principle quantum number of the higher energy level.

The energy levels in hydrogen converge and coalesce:

Energy-level diagram for hydrogen



Since the electron is in the n_1=1 ground state we need to consider series 1. These transitions occur in the UV - part of the spectrum and are known as The Lyman Series.

You can see that as the value of n_2 increases then the value of $1/n_2^2$ decreases. At higher and higher values the expression tends to zero until at n=∞ we can consider the electron to have left the atom resulting in an ion.

The Rydberg expression now becomes:

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - 0\right)$$

Since n_1=1 this becomes:

$$\frac{1}{\lambda} = R = 1.097 * 10^7 \Rightarrow \lambda = 9.116 * 10^{-8}\ (m)$$

We can now find the frequency and hence the corresponding energy:

$$c = \nu * \lambda \Rightarrow \nu = c/\lambda$$

$$\nu = \frac{c}{\lambda} = \frac{3 * 10^8}{9.116 * 10^{-8}} = 3.291 * 10^{15}\ (s^{-1})$$

Now we can use the Planck expression:

$$E = h * v$$

$$E = 6.626 * 10^{-34} * 3.291 * 10^{15} = 2.18 * 10^{-18}\ (J)$$

This is the energy needed to remove 1 electron from 1 hydrogen atom. To find the energy required to ionize 1 mole of H atoms we multiply by the Avogadro Constant:

$$E = 2.18 * 10^{-18} * 6.02 * 10^{23} = 13.123 * 10^5 \left(\frac{J}{mol}\right) = 1312.3\ (\frac{kJ}{mol})$$

Answer: E = 1312.3 kJ*mol^{-1}