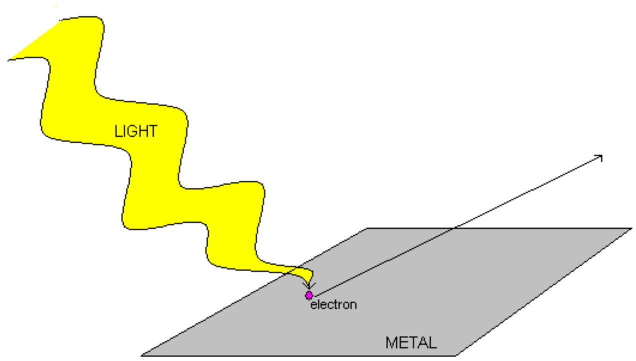
The Photoelectric Effect

1. What is the photoelectric effect?

The photoelectric effect is a phenomenon that Einstein won a Nobel Prize for. What Einstein showed, was that electrons could be emitted from the surface of a metal when light (at a minimum frequency) struck the metal surface.



2. Four ideas to know about it

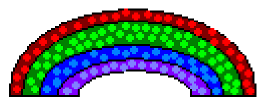
   a. There is a minimum frequency needed to get electrons to emit. An increase in intensity has no effect on electrons if the frequency of the light is below the threshold frequency.

   b. Once the threshold frequency has been reached, an increase in light intensity will cause an increase in the number of electrons emitted.

   c. The greater the frequency of light, the greater the kinetic energy of the emitted electron(s).

   d. Demonstrated that light has a duality in nature. Exhibiting both particulate and energy properties. The only way that these electrons could be knocked out is if the light has a particulate property to it that was able of "bumping" the electrons out.

3. The big discovery of the photoelectric effect was the existence of photons. Photons are little packets of energy contained in light… one very simplistic way to view it is as though it were a skittles rainbow:



.

4. Equation for the energy of a photon

$$E_{photon} = h\nu = \frac{hc}{\lambda}$$

Where h = 6.626 x 10^{-34} J s (Planck's constant)

5. Equation for kinetic energy of an electron

$$KE_{electron} = h\nu - h\nu_0 = E - E_o$$

Where, ν = incident frequency (meaning the frequency of light that you actually used) and ν_0 = threshold frequency (meaning the minimum energy required to eject electron).

6. How do E, ν, and λ relate?

$\uparrow E \uparrow \nu$
$\uparrow E \downarrow \lambda$

7. Microwave radiation has a wavelength on the order of 1.0 cm. Calculate the frequency and the energy of a single photon of this radiation. Calculate the energy of an Avogadro's number of photons of the electromagnetic

radiation.

$$\lambda = 1.0\ \text{cm}\ \frac{\text{m}}{100\ \text{cm}} = 0.010\ \text{m}$$

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\ \text{J s})(3.0 \times 10^{8}\ \text{m/s})}{0.010\ \text{m}} = 1.988 \times 10^{-23}\ \text{J}$$

This value that we have obtained is the amount of energy per photon of light at this frequency.

In order to get an Avogadro's number of photons (or mole of photons), we will have to multiply the value obtained by 6.022×10^{23}.

$$1.988 \times 10^{-23}\ \frac{\text{J}}{\text{photon}}\ \frac{6.022 \times 10^{23}\ \text{photons}}{\text{mol}} = 11.97\ \frac{\text{J}}{\text{mol}}$$

8. The work function of an element is the energy required to remove an electron from the surface of the solid. The work function for lithium is 279.7 kJ/mol. What is the maximum wavelength of light that can remove an electron from an atom in lithium?

"work function" = E_0 (the energy required to emit a mol of electrons) = 279.7 kJ/mole e^-

We are trying to figure out the wavelength associated with emitting a single electron from the lithium metal. The first thing we need to do is determine how much energy is required per electron:

$$\frac{279.7\ \text{kJ}}{\text{mol } e^-}\ \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{electrons}}\ \frac{1000\text{J}}{1\ \text{kJ}} = 4.64 \times 10^{-19}\ \text{J}$$

Now that we know the amount of energy per electron we can solve for λ using:

$$E = \frac{hc}{\lambda} \rightarrow \lambda = \frac{hc}{E}$$

$$\lambda = \frac{(6.626 \times 10^{-34}\ J\ s)(2.998 \times 10^{8}\ \frac{m}{s})}{(4.64 \times 10^{-19}\ J)} = 4.28 \times 10^{-7}\ m = 428\ nm$$

9. It takes 208.4kJ f energy to remove one mole of electrons from the atoms on the surface of rubidium metal. If rubidium metal is irradiated with 254-nm light, what is the maximum kinetic energy the released electron can have?

For this you will need to use the equation:

$$E_{electron} = E - E_0$$

Once again, make sure that you note that the question is asking the amount of energy *an* ejected electron will have (not a mole of ejected electrons). This means that you will have to convert your units for the threshold energy from kJ/mol to J/electron.

$$E_0 = \frac{208.4\ kJ}{mol\ e^-} \cdot \frac{1\ mol}{6.022 \times 10^{23}\ electrons} \cdot \frac{1000 J}{kJ} = 3.46 \times 10^{-19}\ J$$

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\ Js)(2.998 \times 10^{8}\ \frac{m}{s})}{(254 \times 10^{-9}\ m)} = 7.83 \times 10^{-19}\ J$$

$$E_{electron} = 7.83 \times 10^{-19} - 3.46 \times 10^{-19} = \boxed{4.37 \times 10^{-19}\ J}$$