Light and Matter

Chemistry 35
Fall 2000

The Electronic Structure of Atoms

- First, we must consider the properties of

<u>Electromagnetic Radiation (EMR)</u>

Why? (flame emission movie)

•So, just what is **EMR?**

- an *oscillating* electric and magnetic field which travels through space

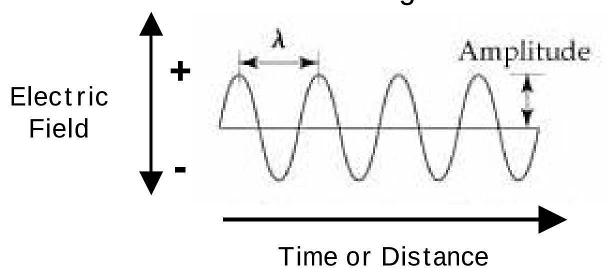
- a discrete series of “particles” that possess a specific energy but have no mass

BOTH!

2

It's a Wave!

- Consider an oscillating electric field:



Characterized by: λ - wavelength

a - amplitude

ν - frequency

3

Wave Properties of EMR

- The product of λ and ν is constant:

$$\boxed{\lambda \times \nu = c}$$

Since ν has units of sec^{-1} and λ has units of length, their product, c, is the *velocity* of the wave:

-in a vacuum, *all EMR* travels at a velocity of:

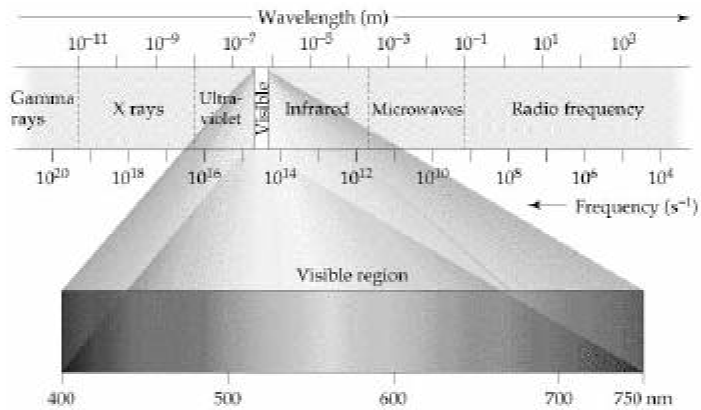
2.99792458 x 10^8 m/s (= c)

("The Speed of Light")

4

The Electromagnetic Spectrum

- The range of wavelengths and frequencies of EMR is extraordinary:



5

Units of Wavelength

- The units used to express wavelength depend upon the region of the electomagnetic spectrum:

TABLE 6.1 Common Wavelength Units for Electromagnetic Radiation

Unit	Symbol	Length (m)	Type of Radiation
Angstrom	Å	10^{-10}	X ray
Nanometer	nm	10^{-9}	Ultraviolet, visible
Micrometer	μm	10^{-6}	Infrared
Millimeter	mm	10^{-3}	Infrared
Centimeter	cm	10^{-2}	Microwave
Meter	m	1	TV, radio

6

Sample Problem

- Microwave ovens use EMR at a frequency of *2.45 GHz* to cook food. What is the **wavelength** of this radiation?

$c = \lambda \nu \rightarrow \lambda = c/\nu$

$$\lambda = \frac{2.9979 \times 10^{8}\ m/s}{2.45 \times 10^{9}\ s^{-1}}$$

$\lambda = 1.21592 \times 10^{-1}$ m

$\boldsymbol{\lambda = 0.122}$ **m**

7

Planck's Quantum Theory

- In 1900, German Physicist ***Max Planck*** proposed:

"Radiant energy may only be absorbed or emitted in discrete amounts: quanta."

-The *energy* of each quantum could then be related to the *frequency* of the EMR:

$$\mathbf{E = h\nu}$$

where: h = $6.6260755 \times 10^{-34}$ J-s (Planck's Constant)

So, the *energy* of a quantum *increases* with: -incr. ν
-decr. λ

8

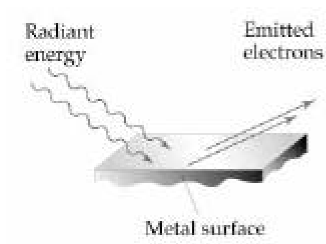
The Photoelectric Effect

- In 1905, Einstein applied this quantum theory to explain the photolectric effect:

-if EMR was absorbed as a *wave*, then the ***number*** of electrons ejected and the ***energy*** of the electrons ejected should vary only with the intensity of the light

•NUMBER of e^-: does vary with EMR intensity

•ENERGY of e^-: vary only with EMR *frequency*



AND: ***no effect*** if freq is below a threshold value!

9

Einstein's Explanation

- View EMR as a collection of *particles* (called **photons**), with each photon having the following energy:

$$\mathbf{E = h\nu}$$

  - Each photon will cause an electron to be *ejected* **IF** the energy of the photon is above a minimum (threshold) value (called the "work function" of the metal)
  - Any energy of the photon *above* that needed to eject the electron would be transferred to the electron as *kinetic energy*
  - Increased EMR intensity translates to an increase in the number of photons (thereby increasing the number of electrons ejected)

10

Photon Energies

- Calculate the *energy* of one photon of light at a wavelength of **700 nm**.

$$E = h\nu \quad \textit{and} \quad c = \lambda\nu$$

So: $\nu = c/\lambda$

Substituting: $\mathbf{E = hc/\lambda}$

$$E = \frac{(6.626 \times 10^{-34}\ \text{J-s})(2.9979 \times 10^{8}\ \text{m/s})}{(700 \times 10^{-9}\ \text{m})}$$

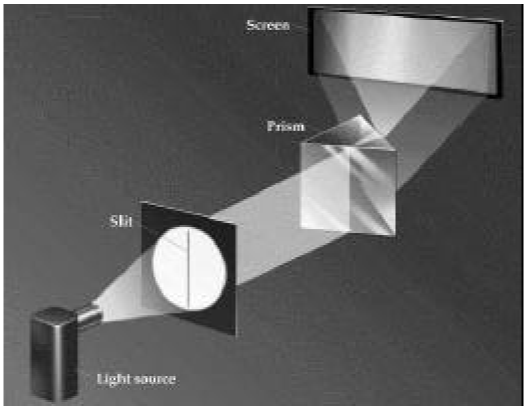
$$= \underline{\boldsymbol{2.84 \times 10^{-19}\ J}}$$

11

Separating Light into Colors

- To figure out what *wavelengths* of light are emitted from a light source, we need to obtain a **spectrum** of the light.

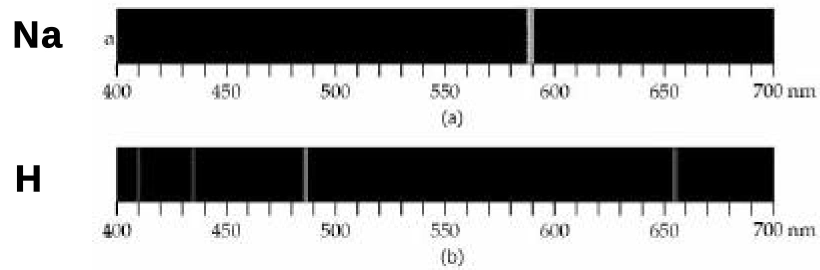
This can be done with a spectrograph:



12

Line Spectra

- The light emitted by elements heated in a flame or electric discharge looks like this:



Line Spectra: *light emitted at* ***discrete*** wavelengths

Each element has a unique spectrum!

13