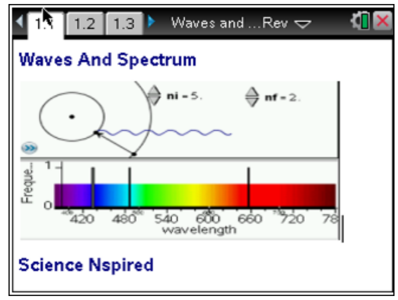
Waves and Spectrum
Student Activity

Name ______________

Class ______________

Open the TI-Nspire document *Waves and Spectrum.tns.*



Why is that neon lights are so attractive? Neon (Ne) is a single type of atom that gives off an orange light to the naked eye. When Neon's light is passed through a spectrum, though, a number of colors appear. Each element will admit a different set of colors called the emission spectrum.

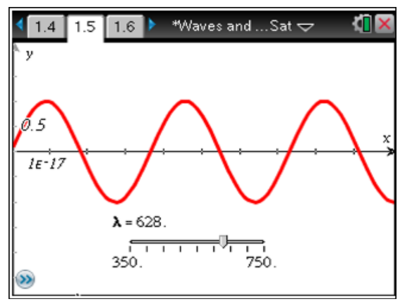
Move to page 1.2.

Press ctrl ▶ and ctrl ◀ to navigate through the lesson.

1. Read the description and instructions on Pages 1.2 through 1.4 of the .tns document.

Move to page 1.5.

2. Adjust the slider, and observe the effect to the wave as well as to the color of the wave.



Move to page 1.6.

Q1. Which wavelength of light has the highest frequency?

- A. Red
- B. Orange
- C. Yellow
- D. Green
- E. Violet

Q2. Calculate the frequency of light with a wavelength of 400nm.

- A. $7.49 \times 10^{5}\ s^{-1}$
- B. $7.49\ s^{-1}$
- C. $7.49 \times 10^{-4}\ s^{-1}$
- D. $7.49 \times 10^{14}\ s^{-1}$

Q3. Calculate the wavelength of light in nm if the frequency is $1.5\text{x}10^{18}$ Hz.

A. $2.0\text{x}10^{-10}$ nm

B. $2.0\text{x}10^{-1}$ nm

C. 2.0 nm

D. $2.0\text{x}10^{10}$ nm

Move to page 1.9.

Light is released in small particles called photons. Photons have wavelengths that help determine their colors. Each color, or wavelength, of light also contains a certain amount of energy. The energy can be determined by calculating the frequency of the photon and multiplying that frequency by Planck's constant (h).

Q4. Which color of light has the highest energy?

A. Red

B. Orange

C. Yellow

D. Green

E. Violet

Q5. Calculate the energy for red light with a frequency of $4.15\text{x}10^{14}$ Hz.

A. $1.60\text{x}10^{-48}$ J

B. $2.75\text{x}10^{-19}$ J

C. $4.78\text{x}10^{-40}$ J

D. $1.39\text{x}10^{6}$ J

Q6. How much energy is released from a photon with a wavelength of 555 nm?

A. $3.68\text{x}10^{-31}$ J

B. $3.68\text{x}10^{-40}$ J

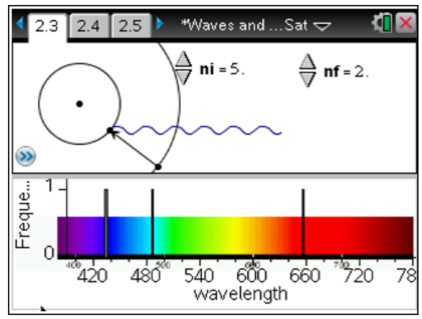
C. $3.58\ \text{x}10^{-19}$ J

D. $3.58\text{x}10^{-28}$ J

Q7. What is the frequency of a photon that releases 4.00×10^{-21} J?

Move to page 2.1.

The physicist Niels Bohr predicted that electrons could only exist at certain energy levels. He determined that the energy levels (n) are indicated by whole numbers (1, 2, 3, 4, ...). He also calculated that when electrons move from a higher energy level to a lower energy level, the energy is released in the form of **light**.



On Page 2.3, *ni* is the initial energy level where the electron starts, and *nf* is the energy level where the electron lands.

- Press ctrl . to capture the wavelength of light on the spectrum.
- Capture *ni* until the value is 10.
- Change *nf* to 2, and repeat until *ni* is 10 again.

Note: The only wavelengths that will appear within the range of 400 nm to 700 nm are visible light wavelengths.

- Change the value of *ni*, and press ctrl . to capture each of the electron drops.

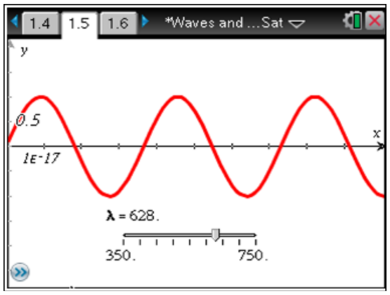
Q8. What electron movement produce visible light?

Q9. Calculate the energy for the 4 wavelengths of light generated in the Balmer series.

Move to page 2.6.

- **Lyman series** are the wavelengths of light that are produced when electrons drop to the first energy level.
- **Balmer series** are the wavelengths of light that are produced when electrons drop to the second energy level.
- **Paschen series** are the wavelengths of light are produced when electrons drop to the third energy level.

3. By changing the *x*-scale on the spectrum page, you can see the Lyman and the Paschen series.



4. Change the 400 nm to 0 nm.

- Put the *nf* to 1, adjust *ni*, and capture the wavelength.
- Observe where the spectrum lines are appearing.

Move to page 2.8.

Q10. Calculate the ΔE for an electron moving from n = 4 to n = 3.

A. 1.06×10^{-19}J
B. -1.06×10^{-19}J
C. 2.42×10^{-19}J
D. 1.36×10^{-19}J