Let’s transition to light and PHOTONS! We will relate this soon to Electrons ;)

Light is defined as both a WAVE and a PARTICLE!

As a PARTICLE, light exists in defined quantities known as Light Quanta or Photons, and these photons have energy associated with them (same is true for ANY moving object – think physics!). A photon is considered to be massless with no electric charge.

As a WAVE, light has properties of waves, including frequency (how often the cycle occurs) and wavelength (the length of each cycle). Mathematically, it looks like this:

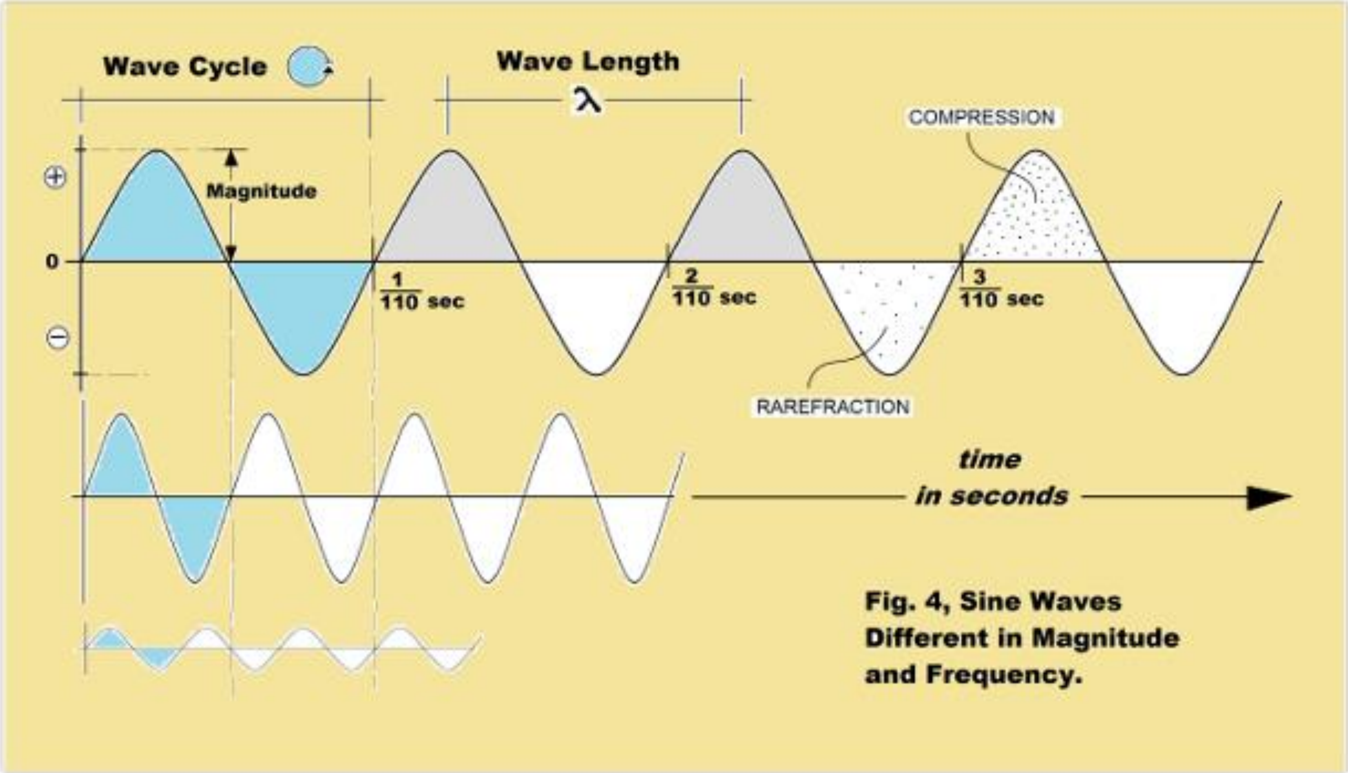


Fig. 4, Sine Waves Different in Magnitude and Frequency.

Because light travels and has energy, we can calculate it's energy and properties!

E = hf

E = Energy of the Photon (Joules)
h = Planck's Constant = 6.626×10^{-34} J-s = **6.6×10^{-34} J-s**
f = frequency of photon (hertz, cycles/second = 1/s))

BUT, for waves, we can relate the frequency to the wavelength via it's speed (true for any constant wave)!

Speed = frequency (f) x **wavelength (λ)**
(wavelength = λ (lambda), measured in meters)

For light then:
Speed of light = frequency (f) x wavelength (λ)
c = frequency (f) x wavelength (λ)

c = f x λ

Where c = speed of light = **3×10^8 meters/second** – Super fast!

Therefore, combining this equation with the equation above:

$E = hf = hc/\lambda$

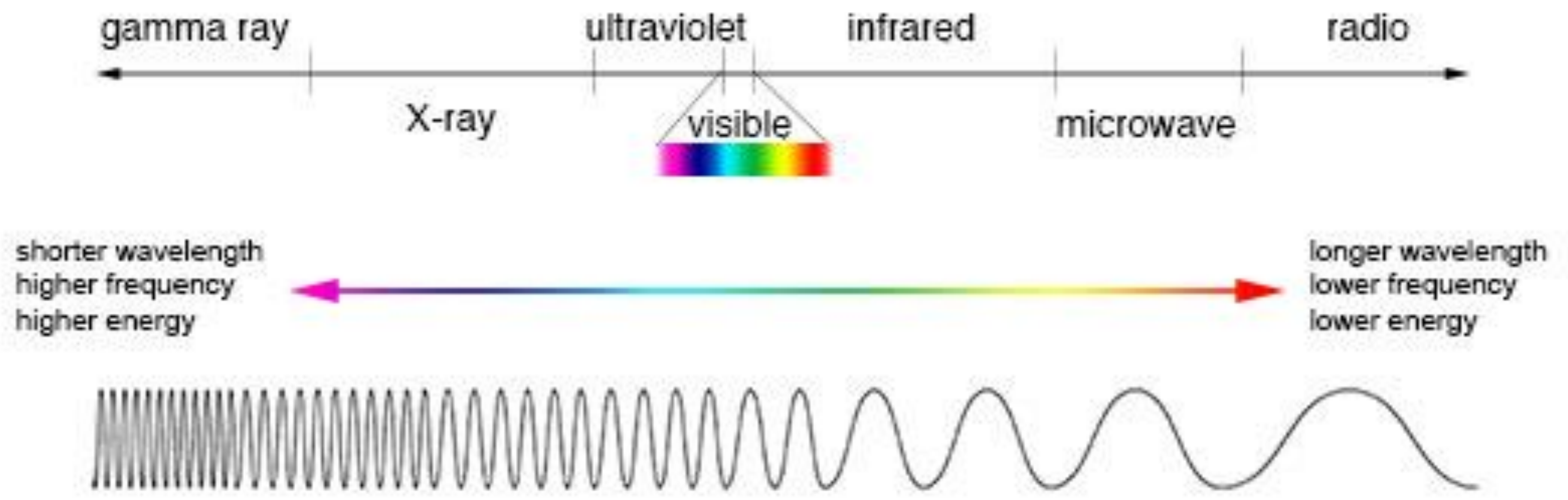
Why does this matter?

1. Properties of different light types can be studied!
2. Fun with exponent math!

$E = hf = hc/\lambda$

Energy and frequency are DIRECTLY related
Energy and wavelength are INVERSELY related

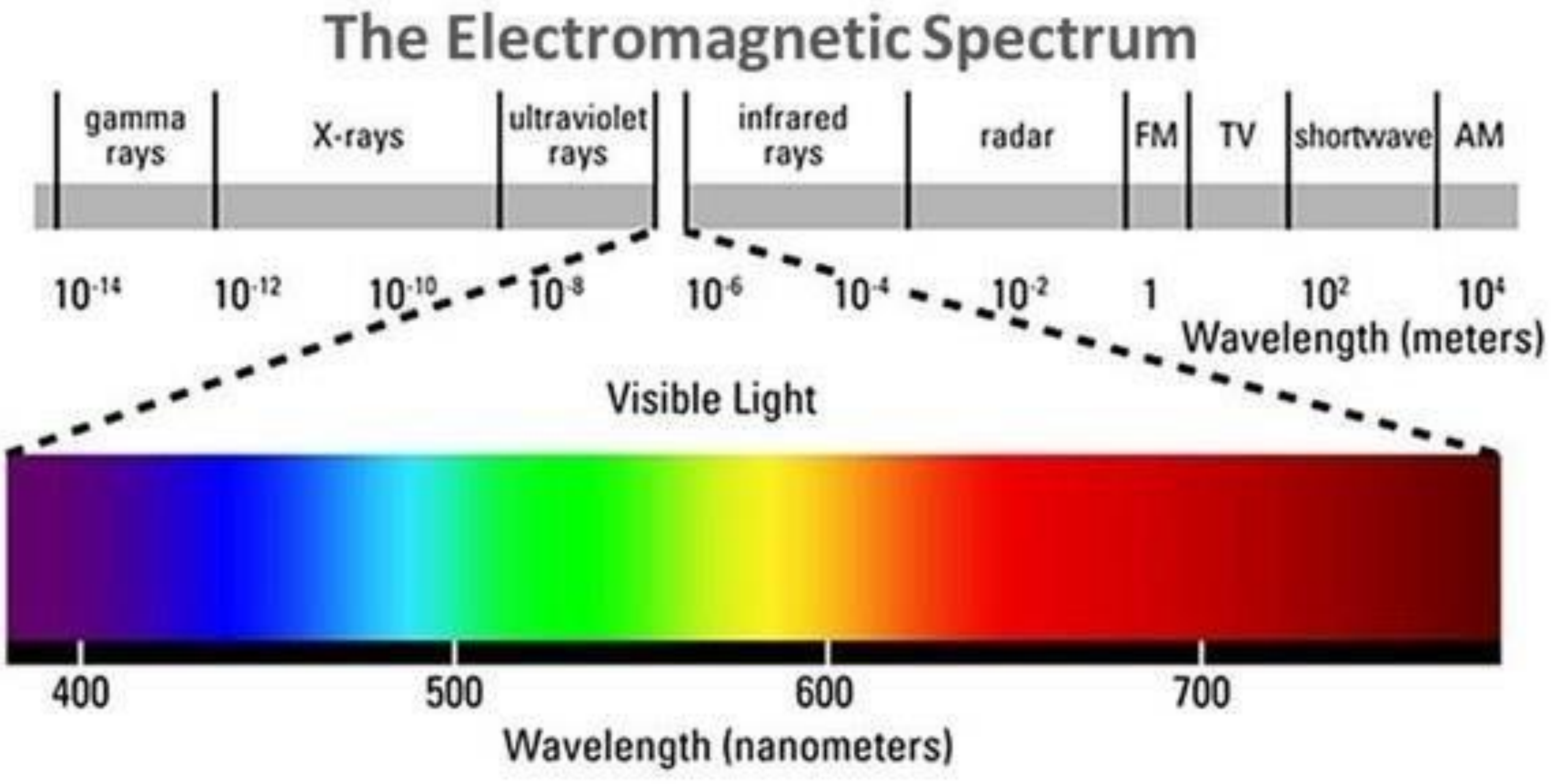
↑ Energy = ↑ frequency = ↓ wavelength
↓ Energy = ↓ frequency = ↑ wavelength



All light types have specified ranges for frequency and wavelength. Commonly, wavelength is used to describe light. The light we see, **visible light**, has wavelengths of 4×10^{-7} meters to 7×10^{-7} meters.

To measure visible light, we normally use nanometers:

1 meter = 1×10^{9} nanometers; Purple = 400 nm; Red = 700nm



Math Problems!

Breakout rooms!

Pick your favorite rainbow color (ROYGBV) from this chart:

Color	Wavelength	Frequency	Energy
Violet	400nm		
Blue	450nm		
Green	500nm		
Yellow	550nm		
Orange	600nm		
Red	700nm		

Calculate your favorite visible light photon's frequency and Energy. $E = hf = hc/\lambda$
For reference: $h = 6.6 \times 10^{-34}$ J-s; $c = 3 \times 10^{8}$ m/s; 1 m = 1×10^{9} nm

This is a multi-step problem. Try as much mental math instead of using a calculator. Rounding off towards the final steps is acceptable ;)

Answer (1 of 6 possibilities)!

My favorite color is Green!

The 1st step is to convert from nanometers to meters:
Green = 500 nm

500nm x 1m/1 x 10^9 nm = 500 x $10^0/10^9$ = **500 x 10^{-9}** = 5 x 10^{-7} m
OR
500nm x 1m/1 x 10^9 nm = 5 x 10^2 nm x 1m/ 10^9 nm = 5 x 10^2 / 10^9 = 5 x 10^{-7} m

The 2nd step is to use our fun, new equations!

To give us the energy, use **$E = hc/\lambda$**

$E = hc/\lambda$ = (6.6 x 10^{-34} J-s)(3 x 10^8 m/s)/ 5 x 10^{-7} m)

(6.6)(3)/(5) x (10^{-34})(10^8)/(10^{-7})

19.8/5 x (10^{-26})/(10^{-7}) ≈ 4 x 10^{-19}

3.96 x 10^{-19} Joules ≈ 4 x 10^{-19} Joules

Answer (1 of 6 possibilities) Continued . . . .

Then, use your answer with this equation: $E = hf$

$E = hf$

3.96×10^{-19} Joules = $(6.6\text{x } 10^{-34}$ J-s) x f

$f = 3.96 \times 10^{-19}/6.6 \times 10^{-34}$

$f = 6 \times 10^{14}$

Whew!

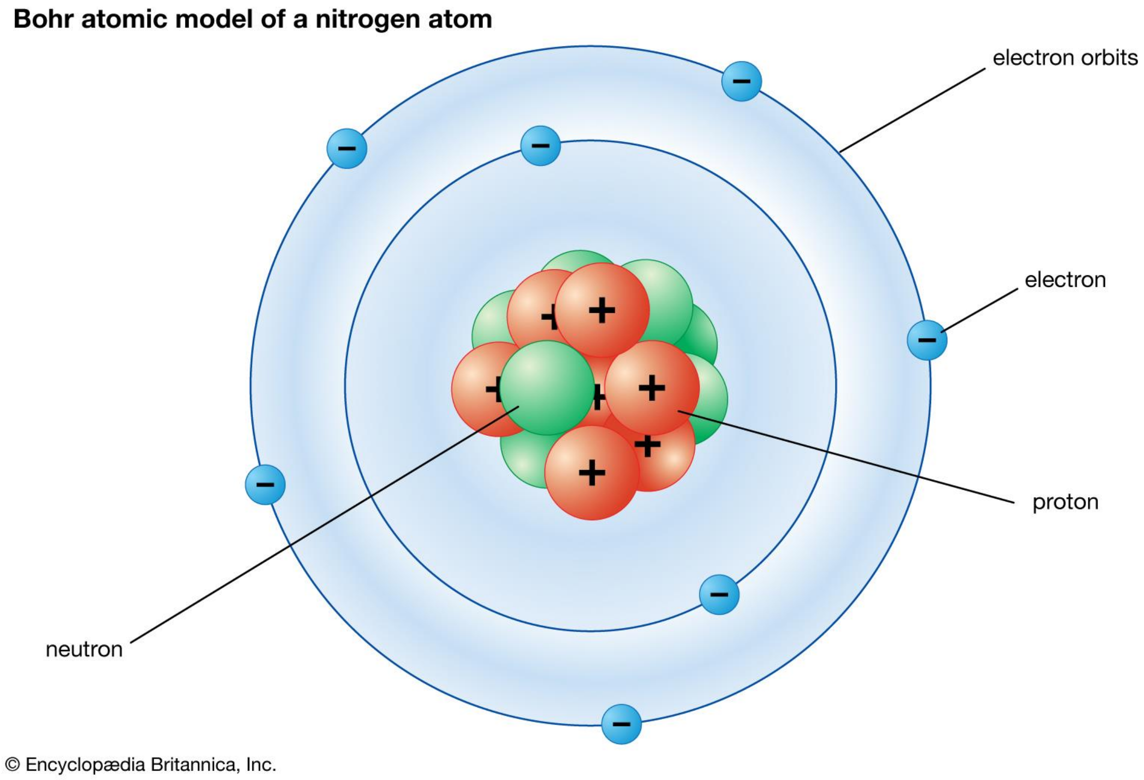
Color	Wavelength	Frequency (1/s)	Energy (J)
Violet	400nm	7.5×10^{14}	4.95×10^{-19}
Blue	450nm	6.67×10^{14}	4.4×10^{-19}
Green	500nm	6×10^{14}	3.96×10^{-19}
Yellow	550nm	5.45×10^{14}	3.6×10^{-19}
Orange	600nm	5×10^{14}	3.3×10^{-19}
Red	700nm	4.29×10^{14}	$2.83\text{x } 10^{-19}$

Let’s go back to the structure of the atom and the electron!

From Slide 1:

Protons & Neutron = exist in nucleus
Electron = exist outside of the nucleus

A *old* version of how to *THINK* about this is called the **Bohr Model of Atom**.
It in essence has electrons orbiting around the nucleus:

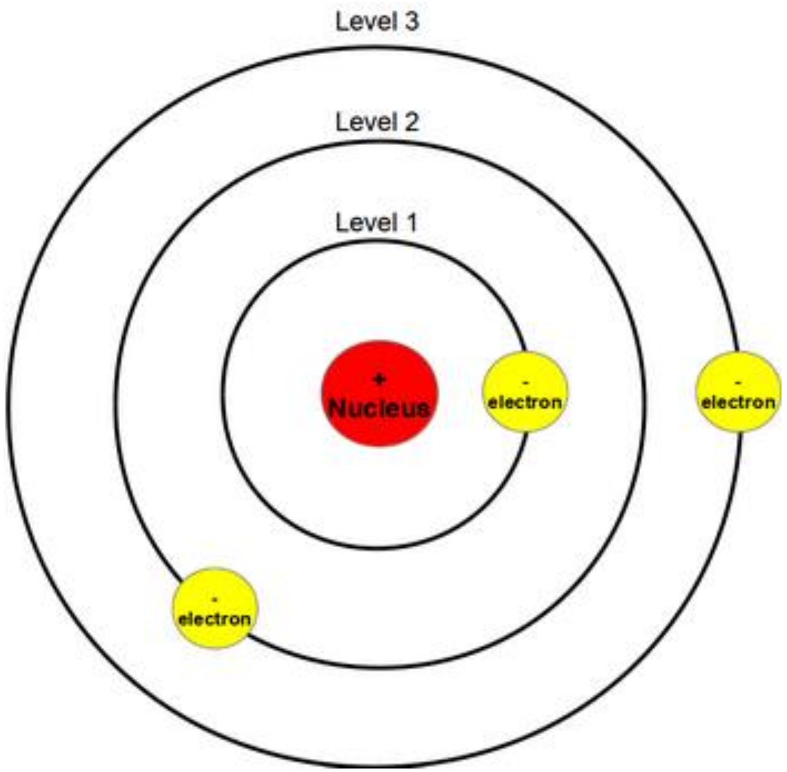


The electron orbits exist in defined energy state or levels (n=1, n=2, n=3, etc.).

What was discovered was that the amount of energy to go from one energy level to another is a specific amount, and this can be related to the energy of photons!

To move to higher energy level (for example, 1→3), an electron **absorbs** a photon (or the energy of a photon) to get to that level.

To move to a lower energy level (for example, 2→1), an electron **emits** energy in the form of a photon.

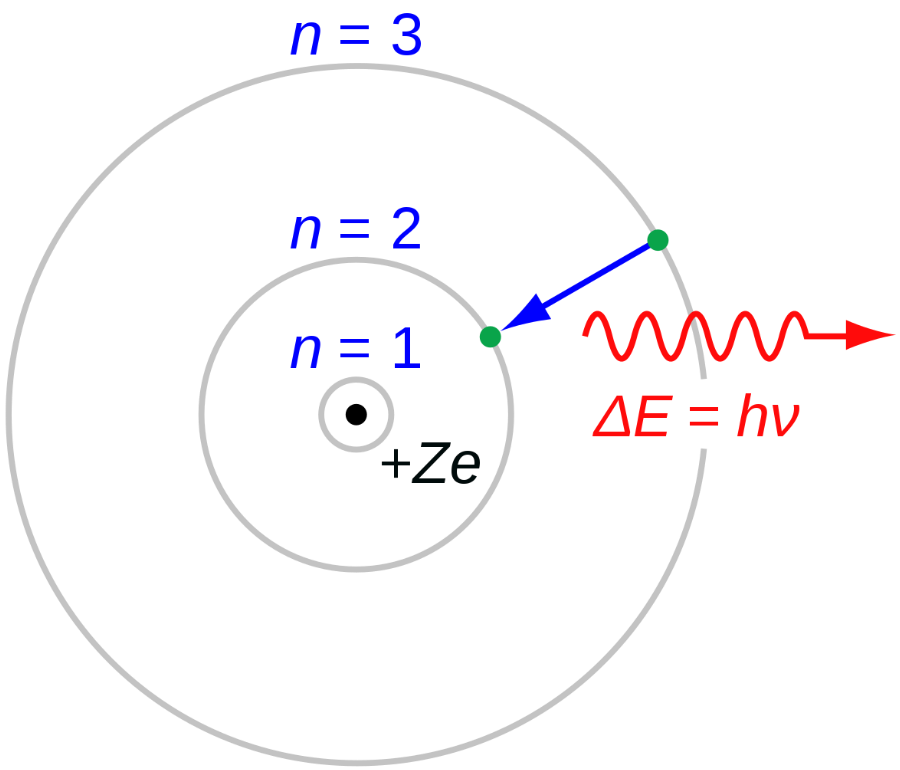


Total Energy Absorbed or Emitted is:

Energy Change = Energy of final level – Energy of initial level

With some amazing physics and math, we get to the Bohr Equation for Energy Transitions for the Hydrogen Atom:

$E = -R_E\,(1/n_f^2 - 1/n_i^2)$, where $R_E = 2.18 \times 10^{-18}$ J, which we'll simplify to 2×10^{-18} J.



Math Problems!

$E = -R_E\,(1/nf^2 - 1/ni^2)$
$R_E = 2 \times 10^{-18}$ J

What energy is absorbed or emitted for the following transitions? No Calculators! Use fraction math and then convert to decimals when necessary

1. An electron moves from the 1st energy level to the 4th energy level.
2. An electron moves from the 4th energy level to the 2nd energy level.
3. An electron moves from the 2nd energy level to the 5th energy level.

Bonus!

1. An electron in the 1st energy leaves the atom.
2. For your answer to #2, what is the closest color of the photon emitted using our previous work? Use calculator for final step only.

Color	Wavelength	Frequency (1/s)	Energy (J)
Violet	400nm	7.5×10^{14}	4.95×10^{-19}
Blue	450nm	6.67×10^{14}	4.4×10^{-19}
Green	500nm	6×10^{14}	3.96×10^{-19}
Yellow	550nm	5.45×10^{14}	3.6×10^{-19}
Orange	600nm	5×10^{14}	3.3×10^{-19}
Red	700nm	4.29×10^{14}	2.83×10^{-19}

Answers!

$E = -R_E (1/nf^2 – 1/ni^2)$
$R_E = 2 \times 10^{-18}$ J

What energy is absorbed or emitted for the following transitions?

1. An electron moves from the 1st energy level to the 4th energy level.

$E = -R_E (1/nf^2 – 1/ni^2) = (-2 \times 10^{-18})(1/4^2 – 1/1^2) = (-2 \times 10^{-18})(1/16 – 1/1) = (-2 \times 10^{-18})(1/16 – 16/16) =$
$(-2 \times 10^{-18})(-15/16) = (2)(15/16)(10^{-18}) = (15/8)(10^{-18}) =$ **1.875×10^{-18} J Absorbed**

2. An electron moves from the 4th energy level to the 2nd energy level.

$E = -R_E (1/nf^2 – 1/ni^2) = (-2 \times 10^{-18})(1/2^2 – 1/4^2) = (-2 \times 10^{-18})(1/4 – 1/16) = (-2 \times 10^{-18})(4/16 – 1/16) =$
$(-2 \times 10^{-18})(3/16) = (-2)(3/16)(10^{-18}) = (-3/8)(10^{-18}) =$ **-0.375×10^{-18} J Emitted**

3. An electron moves from the 2nd energy level to the 5th energy level.

$E = -R_E (1/nf^2 – 1/ni^2) = (-2 \times 10^{-18})(1/5^2 – 1/2^2) = (-2 \times 10^{-18})(1/25 – 1/4) = (-2 \times 10^{-18})(4/100 – 25/100) =$
$(-2 \times 10^{-18})(-21/100) = (2)(21/100)(10^{-18}) = (42/100)(10^{-18}) =$ **0.42×10^{-18} J Absorbed**

Answers . . . . Bonus!

$E = -R_E\,(1/nf^2 - 1/ni^2)$; $R_E = 2 \times 10^{-18}$ J

1. An electron in the 1st energy leaves the atom.

nf = infinity! = ∞

ni = 1

$E = -R_E\,(1/\infty^2 - 1/ni^2) = (-2 \times 10^{-18})(\,1/\,\infty^2 - 1/1^2) = (-2 \times 10^{-18})(\,0 - 1/1) = (-2 \times 10^{-18})(\,-1) = \mathbf{2 \times 10^{-18}}$ **J Absorbed**

2. For your answer to #2, what is the closest color of the photon emitted using our previous work?

Photon Energy = -0.375×10^{-18}

And we know $E = hf$

$0.375 \times 10^{-18} = (6.6 \times 10^{-34})f$

$f = 0.375 \times 10^{-18}/6.6 \times 10^{-34} = \frac{3}{8} \times 10^{-18}/\ 6\,\frac{3}{5} \times 10^{-34} = (\frac{3}{8}\ /\ 6\,\frac{3}{5}\,)(10^{-18}/10^{-34}) = (\frac{3}{8}\,/\frac{33}{5})(10^{16}) = (\frac{3}{8} \times \frac{5}{33})(10^{16}) = (\frac{1}{8} \times \frac{5}{11})(10^{16}) =$
$(\frac{5}{88})(10^{16}) = 0.057 \times 10^{16} = 5.7 \times 10^{14} \rightarrow$ in between Yellow/Green

Color	Wavelength	Frequency (1/s)	Energy (J)
Violet	400nm	7.5×10^{14}	4.95×10^{-19}
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