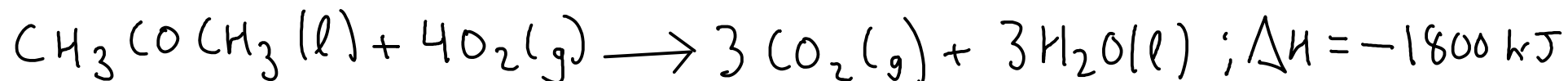


¹⁵⁵ SINCE the enthalpy change does NOT depend on path, this means that we can use standard values for enthalpy to predict the heat change in reactions that we have not tested in a calorimeter.

THERMOCHEMICAL EQUATIONS

- is like a regular chemical equation, except that phase labels are REQUIRED and the enthalpy for the reaction is given along with the equation.



- Why are phase labels required? Because phase changes either absorb or release energy.

$\Delta H = -1800 \text{ kJ}$... what does this mean?

$$1 \text{ mol CH}_3\text{COCH}_3 = -1800 \text{ kJ}$$

$$4 \text{ mol O}_2 = -1800 \text{ kJ}$$

$$3 \text{ mol CO}_2 = -1800 \text{ kJ}$$

$$3 \text{ mol H}_2\text{O} = -1800 \text{ kJ}$$

We treat the enthalpy change as if it's another product of the reaction!



What would be the enthalpy change when 25 g of water are produced by the reaction?

- 1 - Convert mass of water to moles using the formula weight.
- 2 - Convert moles of water to enthalpy change using thermochemical equation

$$18.016 \text{ g H}_2\text{O} = 1 \text{ mol H}_2\text{O} \quad | \quad 3 \text{ mol H}_2\text{O} = -1800 \text{ kJ}$$

$$25 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{-1800 \text{ kJ}}{3 \text{ mol H}_2\text{O}} = \boxed{-830 \text{ kJ}}$$

This reaction is EXOTHERMIC! Energy is released from the reaction to the surroundings. (True for all combustion reactions!)

↪ This equals Q, as long as pressure is constant

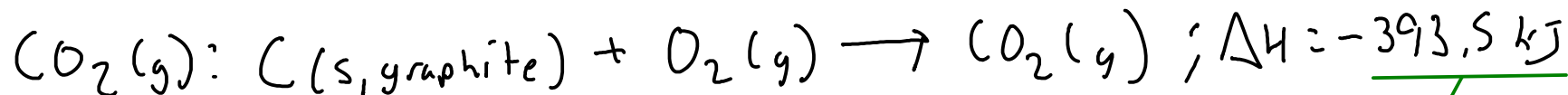
A few more terms related to enthalpy:

- Enthalpy of vaporization / heat of vaporization: The enthalpy change on vaporizing one mole of a substance. (from liquid to vapor)
- Enthalpy of fusion / heat of fusion: The enthalpy change when a mole of liquid changes to the solid state.

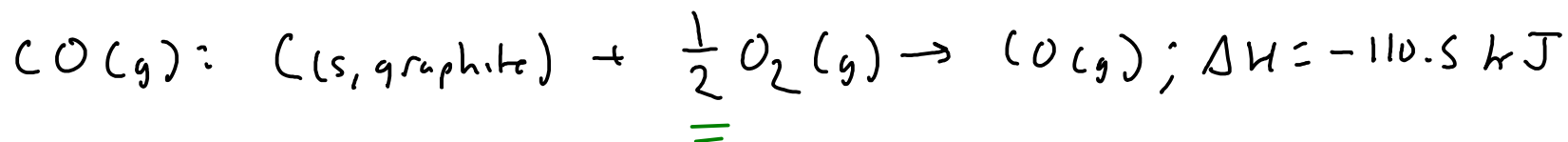
↗ Phase changes require energy, too!

FORMATION REACTIONS

- A reaction that forms exactly one mole of the specified substance from its elements at their STANDARD STATE at 25C and 1 atm pressure.



heat of formation of carbon dioxide ΔH_f° or ΔH_f
 "enthalpy of formation"



you may see fractional coefficients in these formation reactions, because you MUST form exactly one mole of the product!

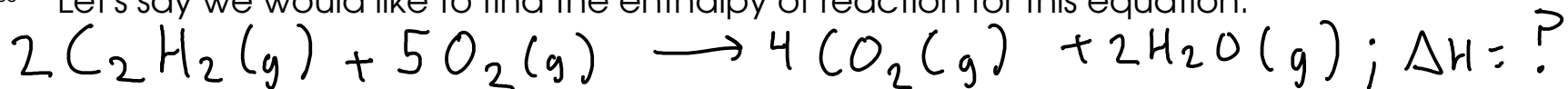
- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.

$$\Delta H_f^\circ, \text{O}_2(g) = 0 \text{ kJ/mol}$$

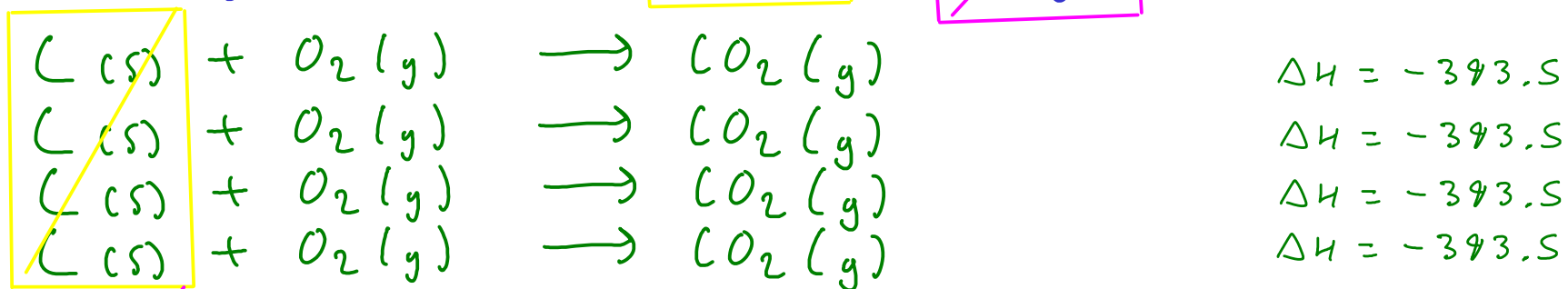
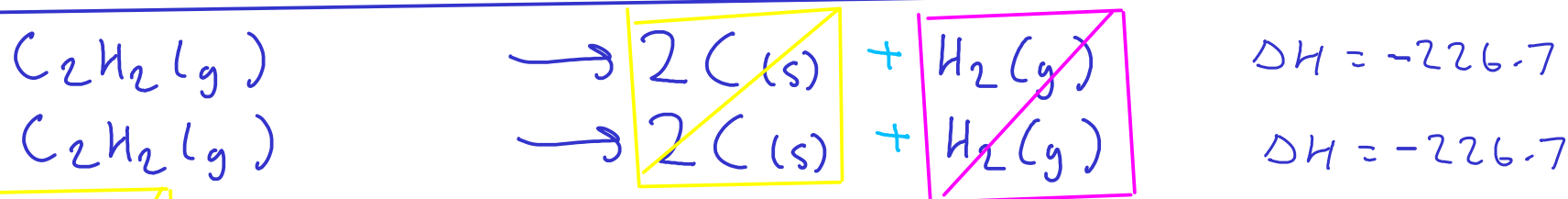
- What are formation reactions good for?

... finding enthalpies for more interesting reactions!

Let's say we would like to find the enthalpy of reaction for this equation:



Hess' Law: If you add two reactions to get a new reaction, their enthalpies also add.



$$\Delta H = 2(-226.7) + 4(-393.5) + 2(-241.8) = \boxed{-2511 \text{ kJ}}$$

Hess' Law using enthalpy of formation:

$$\Delta H = \sum \Delta H_{f, \text{products}} - \sum \Delta H_{f, \text{reactants}}$$

$$\Delta H = \left(\overset{\text{CO}_2}{4(-393.5)} + \overset{\text{H}_2\text{O}}{2(-241.8)} \right) - \left(\overset{\text{C}_2\text{H}_2}{2(226.7)} + \overset{\text{O}_2}{5(0)} \right)$$

$$= \boxed{-2511 \text{ kJ}}$$

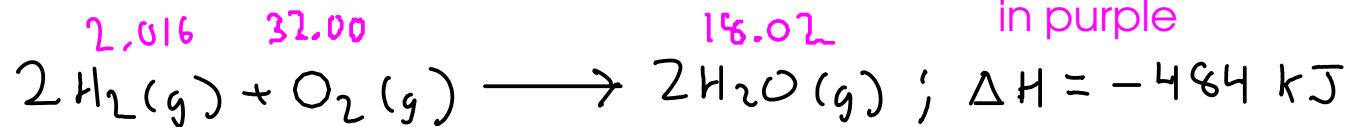
See Appendix C in the textbook for enthalpy of formation data:
p A-8 to A-11

* Remember:

- Multiply each enthalpy by its stoichiometric coefficient from the reaction
- Enthalpy of formation of an element at its standard state is zero
- Watch phase labels. You will usually find SEVERAL enthalpies of formation for a given substance in different phases!
- For ionic substances in solution, remember that they exist as free ions, so look up the aqueous IONS!

160 Example problems:

FORMULA WEIGHTS in g/mol
in purple

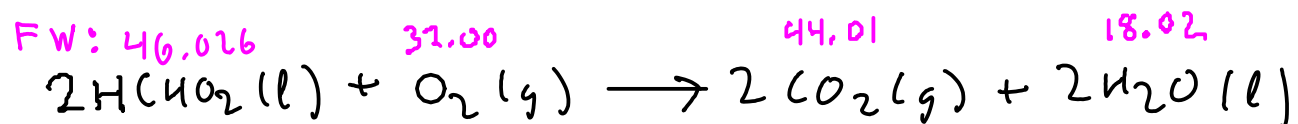


Calculate the enthalpy change for the combustion of 1.00 kg of hydrogen gas.

- 1 - Convert the 1.00 kg of hydrogen to moles (convert to grams, then use formula weight)
- 2 - Convert the moles hydrogen to enthalpy change using thermochemical equation

$$2.016 \text{ g H}_2 = 1 \text{ mol H}_2 \quad | \quad 1 \text{ kg} = 10^3 \text{ g} \quad | \quad 2 \text{ mol H}_2 = -484 \text{ kJ}$$

$$1.00 \text{ kg H}_2 \times \frac{10^3 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{-484 \text{ kJ}}{2 \text{ mol H}_2} = -120000 \text{ kJ per kg H}_2$$



If 50.3 kJ of heat was released when 5.48 g of formic acid are burned at constant pressure, then what is the enthalpy change of this reaction per mole of formic acid?

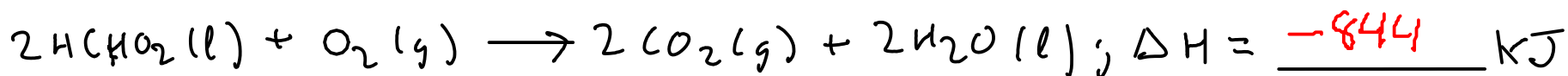
$$Q = -50.3 \text{ kJ} \quad \Delta H = \frac{Q}{\text{mol HCHO}_2}$$

Find moles formic acid:

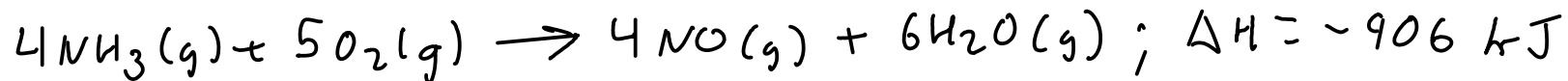
$$5.48 \text{ g HCHO}_2 \times \frac{\text{mol HCHO}_2}{46.026 \text{ g HCHO}_2} = 0.119063 \text{ mol HCHO}_2$$

$$\Delta H = \frac{Q}{\text{mol HCHO}_2} = \frac{-50.3 \text{ kJ}}{0.119063 \text{ mol HCHO}_2} = -422 \frac{\text{kJ}}{\text{mol HCHO}_2}$$

Based on the calculation above, can we complete this thermochemical equation?



We calculated the heat PER MOLE OF FORMIC ACID. This equation is based on the combustion of TWO MOLES of formic acid, so we must double the value.



What is the enthalpy change when 150. L of nitrogen monoxide are formed by this reaction at 25.0 C and 1.50 atm pressure?

1 - Convert 150 L of NO to moles using ideal gas equation

2 - Convert moles of NO to enthalpy using thermochemical equation

$$PV = nRT \quad P = 1.50 \text{ atm} \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

$$\frac{PV}{RT} = n$$

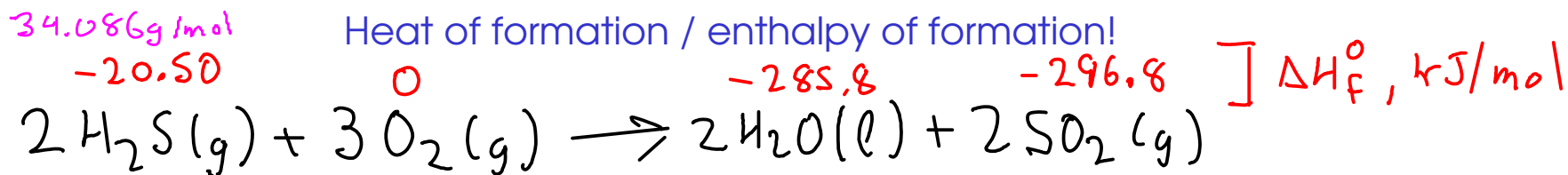
$$V = 150. \text{ L}$$

$$T = 25.0^\circ\text{C} = 298.2 \text{ K}$$

$$\textcircled{1} n_{\text{NO}} = \frac{(1.50 \text{ atm})(150. \text{ L})}{(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298.2 \text{ K})} = 9.19482 \text{ mol NO}$$

$$\textcircled{2} 4 \text{ mol NO} = -906 \text{ kJ}$$

$$9.19482 \text{ mol NO} \times \frac{-906 \text{ kJ}}{4 \text{ mol NO}} = \boxed{-2080 \text{ kJ} = \Delta H}$$



What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?

1 - Find the enthalpy of this reaction with Hess's Law. Use enthalpies of formation.

(See Appendix C, p A-8)

2 - Convert 25.0 g of hydrogen sulfide to moles using formula weight.

3 - Convert moles hydrogen sulfide to enthalpy using thermochemical equation.

$$\textcircled{1} \Delta H = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$= [2(-285.8) + 2(-296.8)] - [2(-20.50) + 3(0)] = -1124.2 \text{ kJ}$$

Thermochemical equation:

... for the reaction as written!

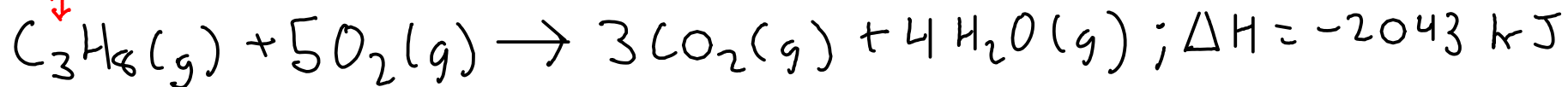


$$34.086 \text{ g H}_2\text{S} = 1 \text{ mol H}_2\text{S} \quad | \quad 2 \text{ mol H}_2\text{S} = -1124.2 \text{ kJ}$$

$$25.0 \text{ g H}_2\text{S} \times \frac{1 \text{ mol H}_2\text{S}}{34.086 \text{ g H}_2\text{S}} \times \frac{-1124.2 \text{ kJ}}{2 \text{ mol H}_2\text{S}} = \boxed{-412 \text{ kJ}}$$

$\textcircled{2}$
 $\textcircled{3}$

propane



Calculate the volume of propane gas at 25.0 C and 1.08 atm required to provide 565 kJ of heat using the reaction above.

- 1 - Convert the energy requirement to moles propane using the thermochemical equation
- 2 - Convert moles propane to volume using the ideal gas equation

$$\textcircled{1} \quad -2043 \text{ kJ} = \text{mol C}_3\text{H}_8$$

$$-565 \text{ kJ} \times \frac{\text{mol C}_3\text{H}_8}{-2043 \text{ kJ}} = 0.276554 \text{ mol C}_3\text{H}_8$$

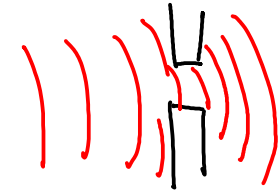
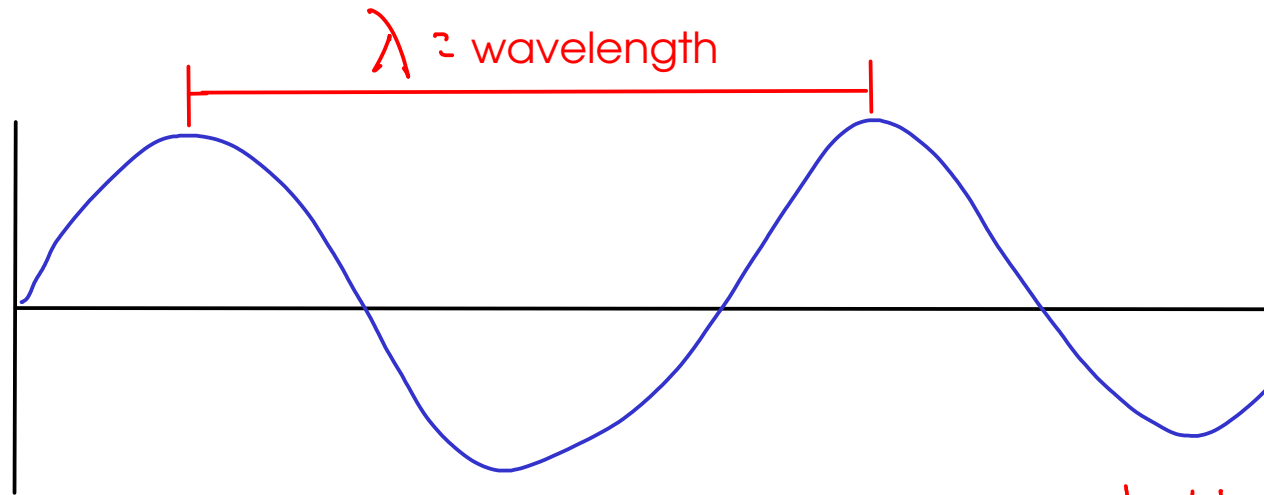
$PV = nRT$ $V = \frac{nRT}{P}$	$n = 0.276554 \text{ mol C}_3\text{H}_8$ $R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$ $T = 25.0^\circ = 298.2 \text{ K}$ $P = 1.08 \text{ atm}$
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$$V = \frac{(0.276554 \text{ mol C}_3\text{H}_8) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298.2 \text{ K})}{(1.08 \text{ atm})}$$

$$= 6.27 \text{ L C}_3\text{H}_8 \text{ needed}$$

END OF CHAPTER 6

LIGHT



Diffraction

$$\text{frequency} = \text{wavelengths} / \text{time} = \nu \quad \text{s}^{-1} : \text{Hertz } \approx, \text{ Hz}$$

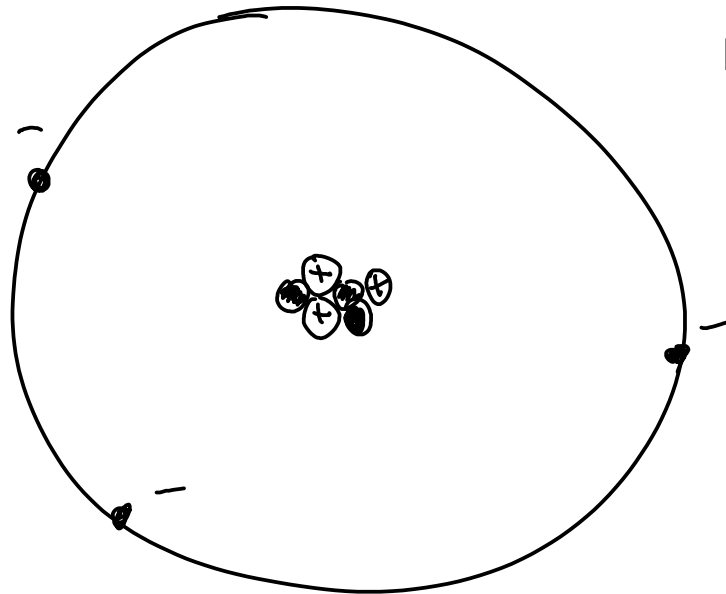
- Light has properties of WAVES such as DIFFRACTION (it bends around small obstructions).
- Einstein noted that viewing light as a particle that carried an energy proportional to the FREQUENCY could explain the PHOTOELECTRIC EFFECT!

$$E_{\text{photon}} = h \nu$$

h — Planck's constant: $6.63 \times 10^{-34} \text{ J}\cdot\text{s}$
 ν — photon = particle or packet of light

(The photoelectric effect is the emission of electrons from a metal caused by exposure to light. Einstein discovered that if the light were not of the correct FREQUENCY, increasing the INTENSITY of the light would not cause electron emission. He concluded that individual photons must have enough energy to excite an electron - i.e. they must have the appropriate frequency.)

The photoelectric effect and Einstein's ideas about the energy content of light led us to discover a new model for the atom! How? Let's start with the nuclear model:



Nuclear model:

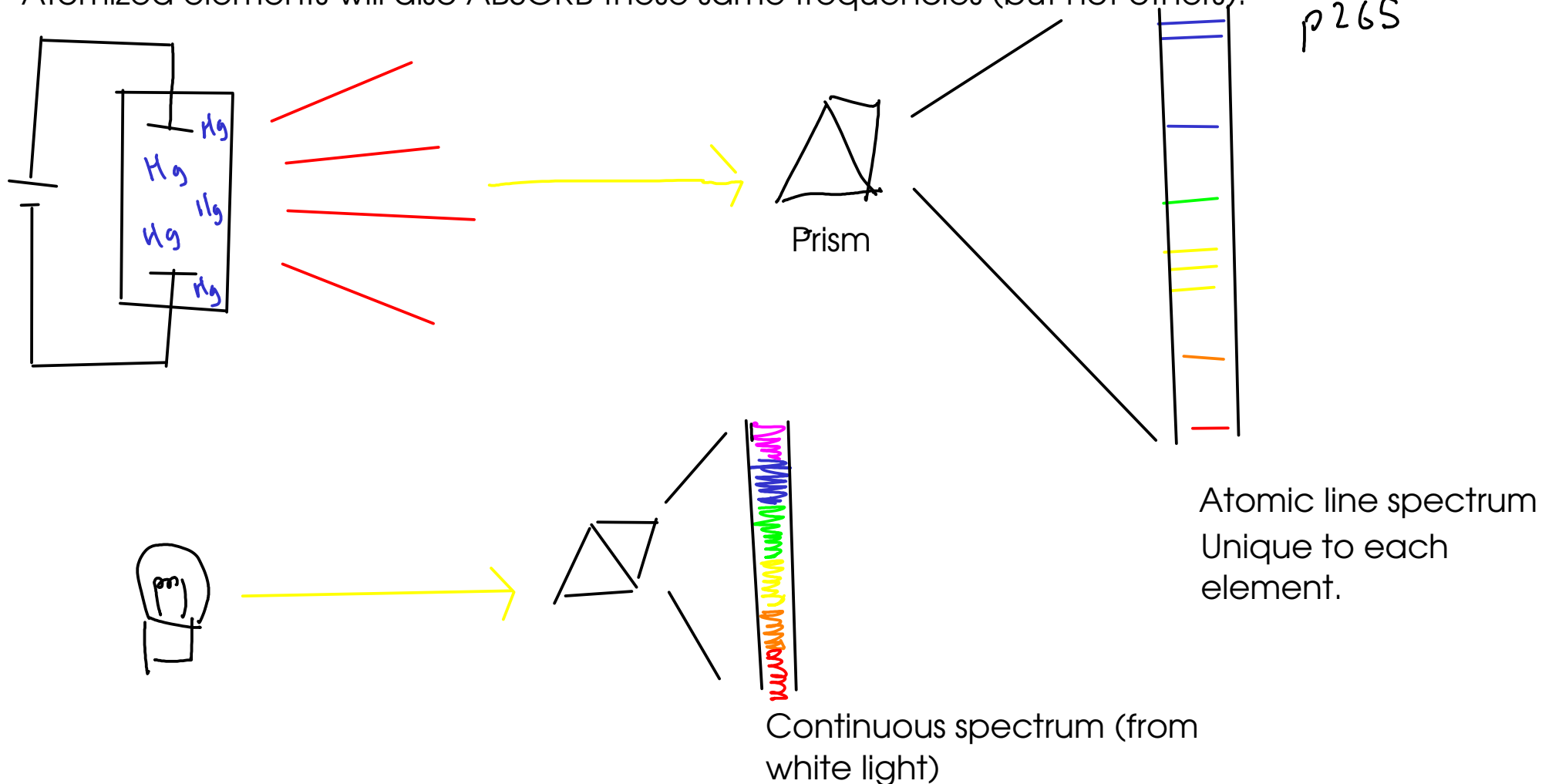
- Protons and neutrons in a dense NUCLEUS at center of atom
- Electrons in a diffuse (mostly empty) ELECTRON CLOUD surrounding NUCLEUS.

... so what's wrong with the nuclear model? Among other things, it doesn't explain ...

ATOMIC LINE SPECTRA

- if you take element and ATOMIZE it, if excited by energy it will emit light at unique frequencies. The set of emitted frequencies is called an ATOMIC LINE SPECTRUM.

- Atomized elements will also ABSORB these same frequencies (but not others)!



... so, why don't atoms by themselves emit continuous spectra like a flashlight would?