

## Sample Questions for Chem 002 Final WS10

### 1. MSDS (the rest listed on review):

- Proper attire – goggles, closed toe shoes, long pants or skirt or lab apron
- Acid Spill – neutralize with sodium bicarbonate
- Bunsen Burners – do not light if flammable biproducts (e.g., H<sub>2</sub> gas) are present
- Phenolphthalein – has a laxative effect when ingested
- Types of radiation (listed below) are stopped by what type of material?
  - alpha – paper or hand
  - beta – aluminum (goes through paper, hand)
  - gamma – lead (goes through paper, hand, aluminum)
  - neutron – concrete (goes through paper, hand, aluminum, lead)

### 2. Radioactive Decay:

- Balance the following radioactive decay equations:



- Determine the specific decay constant, initial activity and half-life of a radioactive isotope. Given

Time, minutes	Counts/Min	ln (Counts/Min)
0		
2	14472	9.58
3	14328	9.57
4	14248	9.56
5	14095	9.55
6	13920	9.54
10	13359	9.50

- Determine the specific decay constant, k, for this radioactive decay.

$$k = -m \quad m = (y_2 - y_1) / (x_2 - x_1) = (9.58 - 9.50) / (2 - 10) \text{ min} = -0.01$$

$$k = 0.01 \text{ min}^{-1}$$

- Determine the initial activity, A<sub>0</sub>.

$$y = mx + b \quad \ln A = -kt + \ln A_0$$

$$9.50 = -(0.01)(10) + \ln A_0$$

$$9.50 = -0.10 + \ln A_0$$

$$9.60 = \ln A_0$$

$$A_0 = e^{\ln A_0} = e^{9.60} = 14764 \text{ counts/min}$$

- Determine the half-life.

$$t_{1/2} = \ln 2 / k = 0.693 / 0.01 \text{ min}^{-1} = 69.3 \text{ min}$$

### 3. Heat of Neutralization:

A reaction of 100mL of 1.35M HCl and 100mL of 1.76M NaOH is monitored and the following temperatures were recorded: starting temperature = 24.6 °C; and final temperature = 38.8 °C. Calculate the  $\Delta H$  of this reaction.

Given:  $C_p$  of solution (J/K) = 4.13 J/(g · K) \* Volume of solution in mL (1 mL  $\approx$  1 g for aqueous soln)  
 $C_p$  of calorimeter (J/K) = 50  
 $Q = (-\text{total } C_p * \Delta T)$   
 $\Delta H = Q/n$   
 $\Delta T = T_f - T_i$   
 $n = \# \text{ of moles reacted}$

a. Determine the change in temperature for the system.

$$\Delta T = T_f - T_i = 38.8 \text{ }^\circ\text{C} - 24.6 \text{ }^\circ\text{C} = 14.2 \text{ }^\circ\text{C}$$

$$\text{Since we calculated } \Delta T, 14.2 \text{ }^\circ\text{C} = 14.2 \text{ K}$$

b. Determine the  $C_p$  of the solution (J/K).

$$C_{p(\text{soln})} = 4.13 \text{ (J/g}\cdot\text{K)} \times (1 \text{ g} / 1 \text{ ml}) \times (200 \text{ ml}) \text{ OR}$$

$$C_{p(\text{soln})} = 4.13 \text{ (J/ml}\cdot\text{K)} \times (200 \text{ ml})$$

$$C_{p(\text{soln})} = 826 \text{ J/K}$$

c. Determine the **total**  $C_p$  of the system.

$$C_{p(\text{sys})} = C_{p(\text{soln})} + C_{p(\text{cal})} =$$

$$C_{p(\text{soln})} = (826 + 50) \text{ J/K} = 876 \text{ J/K}$$

d. Determine the number of moles of the acid and the base. Which is the **limiting reagent**?

$$(1.35 \text{ mole} / \text{L HCl}) \times (100 \text{ ml}) \times (1 \text{ L} / 1000 \text{ ml}) = 0.135 \text{ mole HCl}$$

$$(1.76 \text{ mole} / \text{L NaOH}) \times (100 \text{ ml}) \times (1 \text{ L} / 1000 \text{ ml}) = 0.176 \text{ mole NaOH}$$

Since HCl is the limiting reagent, the moles of solution is 0.135 mole.

e. Determine the **Heat Transfer, Q**, for the reaction.

$$Q = [(-876 \text{ J/K}) (14.2 \text{ K})]$$

$$Q = -12,439.2 \text{ joule}$$

f. Determine the **change in enthalpy,  $\Delta H$** , for the reaction.

$$\Delta H = (-12,439.2 \text{ joule}) / (0.135 \text{ mole soln})$$

$$\Delta H = -92,142 \text{ joule} / \text{mole}$$

**4. Heat of Fusion.** An ice cube with mass 9.53 grams (presume  $T_i = 0\text{ }^\circ\text{C}$ ) is placed in a calorimeter containing 111.24 grams of distilled water at a temperature of  $23.2\text{ }^\circ\text{C}$ . After equilibration, the final temperature was  $15.8\text{ }^\circ\text{C}$ .

Given:  $\Delta H_{\text{total}} = \Delta H_{\text{ice}} + \Delta H_{\text{water}} + \Delta H_{\text{calorimeter}} + \Delta H_{\text{fus}} = 0$

$$\Delta H_{\text{water}} (\text{J}) = C_{p,\text{H}_2\text{O}} * (\text{mass}) * \Delta T \qquad \Delta H_{\text{ice}} (\text{J}) = C_{p,\text{ice}} * (\text{mass}) * \Delta T \qquad H_{\text{calorimeter}} (\text{J}) = C_{p,\text{Cal}} * \Delta T$$

$$C_{p,\text{H}_2\text{O}} = 4.18 \text{ J}/(\text{g}^\circ\text{C}) \qquad C_{p,\text{Cal}} = 50 \text{ J}/^\circ\text{C} \qquad \Delta T = T_f - T_i$$

a. Determine the  $\Delta H_{\text{water}}$ .

$$\begin{aligned} \Delta H_{\text{water}} (\text{J}) &= C_{p,\text{H}_2\text{O}} * (\text{mass}) * \Delta T \\ &= 4.18 \text{ J}/(\text{g}^\circ\text{C}) \times (111.24 \text{ grams}) \times (15.8\text{ }^\circ\text{C} - 23.2\text{ }^\circ\text{C}) \\ &= -3441 \text{ J} \end{aligned}$$

b. Determine the  $\Delta H_{\text{ice}}$ . (*Hint for  $C_p$  – The ice has melted.*)

Since  $C_{p,\text{ice}} = C_{p,\text{H}_2\text{O}}$

$$\begin{aligned} \Delta H_{\text{ice}} (\text{J}) &= C_{p,\text{H}_2\text{O}} * (\text{mass}) * \Delta T \\ &= 4.18 \text{ J}/(\text{g}^\circ\text{C}) \times (9.53 \text{ grams}) \times (15.8\text{ }^\circ\text{C} - 0.0\text{ }^\circ\text{C}) \\ &= 629 \text{ J} \end{aligned}$$

c. Determine the  $\Delta H_{\text{calorimeter}}$ .

$$\begin{aligned} H_{\text{calorimeter}} (\text{J}) &= C_{p,\text{Cal}} * \Delta T \\ &= 50 \text{ J}/^\circ\text{C} \times (15.8\text{ }^\circ\text{C} - 23.2\text{ }^\circ\text{C}) \\ &= -370 \text{ J} \end{aligned}$$

d. Determine the  $\Delta H_{\text{fus}}$  for **one gram** of ice.

(*Hint:* For a calorimeter (i.e., closed systems)  $\Delta H_{\text{total}} = 0$ )

$$\begin{aligned} \Delta H_{\text{total}} &= \Delta H_{\text{ice}} + \Delta H_{\text{water}} + \Delta H_{\text{calorimeter}} + \Delta H_{\text{fus}} = 0 \\ \Delta H_{\text{fus}} &= -(\Delta H_{\text{ice}} + \Delta H_{\text{water}} + \Delta H_{\text{calorimeter}}) \\ &= -(629 \text{ J} - 3441 \text{ J} - 370 \text{ J}) \\ &= -3182 \text{ J} \end{aligned}$$

So  $\Delta H_{\text{fus}}$  for **one gram** of ice

$$\begin{aligned} \Delta H_{\text{fus}} &= -3182 \text{ J} / 9.53 \text{ g} \\ &= 334 \text{ J/g} \end{aligned}$$

f. If heat transfers from the system (solute) to the surroundings (solvent), then  $\Delta H$  is negative ( $\Delta H < 0$ ), and the reaction is defined as (endothermic / **exothermic**) and the temperature of the solvent will go (**up** / down).

g. If heat transfers from the surroundings (solvent) to the system (solute), then  $\Delta H$  is positive ( $\Delta H > 0$ ), and the reaction is defined as (**endothermic** / exothermic) and the temperature of the solvent will go (up / **down**).

h. The heat of neutralization experiment was an (endothermic / **exothermic**) reaction .

i. The heat of fusion experiment was an (**endothermic** / exothermic) reaction.

j. This term means “the techniques that are used to measure enthalpy”: **Calorimetry**

k. This term means “the energy needed to raise the temperature of an object  $1^\circ\text{C}$ ”: **Heat Capacity**

l. This term means “the energy needed to raise the temperature of one gram of a substance  $1^\circ\text{C}$ ”: **Specific Heat**

m. The heat capacity is an extrinsic property. Define intrinsic and extrinsic properties and give an example of each.

Intrinsic properties are inherent properties usually physical. For example, when a piece of wood is cut, each piece still has the appearance of wood.

Extrinsic properties are dependent upon the amount of an object present. For example, when a small piece of wood is burned it generates less heat, than when a large piece of wood is burned.

**5. Antacids:** You are given 1.12 M HCl and 1.48 M NaOH. The antacid you use contains 300 mg of CaCO<sub>3</sub> and 100 mg of Al(OH)<sub>3</sub>. If the antacid dissolved in 35.0 ml of HCl and was then back titrated with 21.8 ml of NaOH, find the following:

- a. The original number mmoles of HCl used to dissolve the antacid and neutralize the base.

$$(1.12 \text{ mmole / ml HCl}) \times (35.0 \text{ ml HCl}) = 39.2 \text{ mmole HCl}$$

- b. The number of mmoles of NaOH used to backtitrate the acid.

$$(1.48 \text{ mmole / ml NaOH}) \times (21.8 \text{ ml NaOH}) = 32.3 \text{ mmole NaOH}$$

- c. The number of mmoles of antacid used to neutralize only the antacid (a.k.a. the excess HCl).

$$\text{Excess HCl} = \text{mmole HCl} - \text{mmole NaOH} = 39.2 - 32.3 = 6.9 \text{ mmole HCl}$$

- d. Write the balanced equations for the neutralization of the antacid (Both CaCO<sub>3</sub> and Al(OH)<sub>3</sub> ).



- e. Using the number of mg in the tablet, calculate the mmoles of each component (Both CaCO<sub>3</sub> and Al(OH)<sub>3</sub> ).

$$300 \text{ mg CaCO}_3 \times (1 \text{ mmole} / 100 \text{ mg CaCO}_3) = 3.00 \text{ mmole CaCO}_3$$

$$100 \text{ mg Al(OH)}_3 \times (1 \text{ mmole} / 78 \text{ mg Al(OH)}_3) = 1.30 \text{ mmole Al(OH)}_3$$

- f. Based on the mmoles of each component, calculate the theoretical number of mmoles of HCl that should have been needed to neutralize the antacid. (*Hint: Use the mole ratios.*)

$$3.00 \text{ mmole CaCO}_3 \times (2 \text{ mmole HCl} / 1 \text{ mmole CaCO}_3) = 6.00 \text{ mmole HCl}$$

$$1.30 \text{ mmole Al(OH)}_3 \times (3 \text{ mmole HCl} / 1 \text{ mmole Al(OH)}_3) = 3.90 \text{ mmole HCl}$$

- g. What was the total number of theoretical mmoles of HCl that should have been neutralized?

$$6.00 + 3.90 \text{ mmole HCl} = 9.9 \text{ mmole HCl}$$

- h. Compare the theoretical (g.) to the actual (c.). What are possible reasons this discrepancy could have occurred?

$$\text{Actual is } 6.9 \text{ mmole} < \text{Theoretical } 9.9 \text{ mmole}$$

**Possible Reasons:**

Student may not have performed the titration accurately.

Manufacturer may not have quality control standards that ensure the amount of ingredients.

Binders and other additives may have interfered with the effectiveness of the antacids.

**6. Spectrophotometry:** Using a Spectrophotometer (Spec 20), a student recorded below the Percent Transmittance data for the following solutions:

Red Dye Standard (6.30 ppm)  
Blue Dye Standard (5.05 ppm)  
Purple Unknown

	400 nm	450 nm	500 nm	550 nm	600 nm	650 nm
Red Std	63.5	48.5	23.5	38.6	78.3	98.5
Blue Std	80.5	99.0	82.5	56.5	8.5	72.4
Purple Unk	79.3	72.5	35.5	85.8	45.5	65.3

a. Calculate the Absorbance for each of the %T listed above .

	400 nm	450 nm	500 nm	550 nm	600 nm	650 nm
Red Std	$\log(100/63.5)$ 0.197	0.314	0.629	0.413	0.106	0.007
Blue Std	0.094	0.004	0.084	0.248	1.07	0.140
Purple Unk	0.101	0.140	0.450	0.067	0.342	0.185

b. Determine the following from the data calculated in Part 1 (2 pts):

Red Dye Max. Absorbance = 0.629 at 500 nm ( $\lambda$  Max)

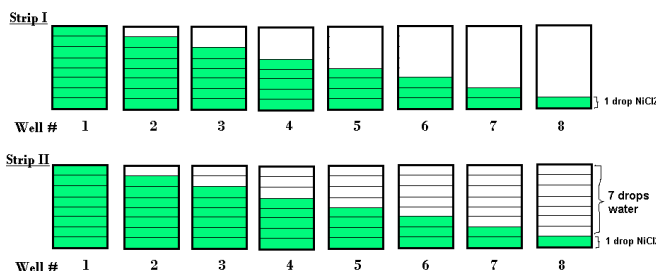
Blue Dye Max. Absorbance = 1.07 at 600 nm ( $\lambda$  Max)

c. Calculate the Absorbance Ratio of the Unknown/Standard at ( $\lambda$  Max).

d. Calculate the Dye Concentration in the Unknown. (Standard Concentrations given above.)

	Abs of Unknown (at $\lambda$ Max)	Abs of Standard (at $\lambda$ Max)	Abs Ratio Unk/Std (at $\lambda$ Max)	Dye Conc. in Unknown
Red in Purple	0.450	0.629	$(0.450/0.629) = 0.715$	$(0.715 \times 6.30) = 4.50$
Blue in Purple	0.342	1.07	$(0.342/1.07) = 0.320$	$(0.320 \times 5.05) = 1.62$

**7. Colorimetry:** Using the well strips below, the student put the following number of drops in the wells. In strips I& II, 1-8 drops of blue dye standard solution (5.05 ppm) were added as shown in the diagram. In strip II, additional drops of water were added in order to have the same total volume of 8 drops for each well.



**Given:** The student found that the unknown solution of blue dye matched well #7 on Strip II.

- What is changing in the first well strip – concentration or **pathlength**?
- What is changing in the second well strip – **concentration** or pathlength?
- Looking from the top how does the intensity compare for Strip 1 to Strip 2?  
more intense – **the same** – less intense
- Using  $C_1V_1 = C_2V_2$ , what is the approximate concentration in ppm for the unknown?  
In Well #7  $5.05\text{ppm} (2 \text{ drops}) = C_2 (8 \text{ drops})$   
 $1.26(25) \text{ ppm} = C_2$

## 8. Equilibrium Constant & Beer's Law:

1. Why is it important that you avoid any loss of solutions when transferring and measuring solutions in an equilibrium experiment similar to the one done in class?

All of the concentrations are calculated on the basis of volumes of the solutions being additive and the total solution present. If any of the solution is lost due to spillage, the calculated values of concentration would be less than actual:

i.e.,  $x \text{ moles} / y \text{ volume} < x \text{ moles} / (y-z) \text{ volume}$ .

e.g.,  $1 \text{ mole} / 0.4 \text{ L} < 1 \text{ mole} / 0.39 \text{ L}$

2. According to the Beer's Law equation,  $A = abc$ . Why is it you did not need to know the value for the molar absorptivity constant or the pathlength for the calculations you did in this experiment?

At any given wavelength, the molar absorptivity is constant. The pathlength is also constant. So the concentration of the thiocyanatoiron (III) ion is then directly proportional to the absorbance.

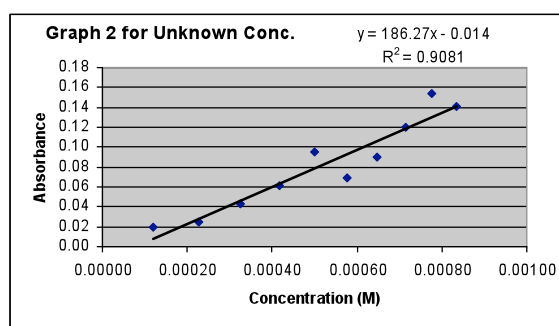
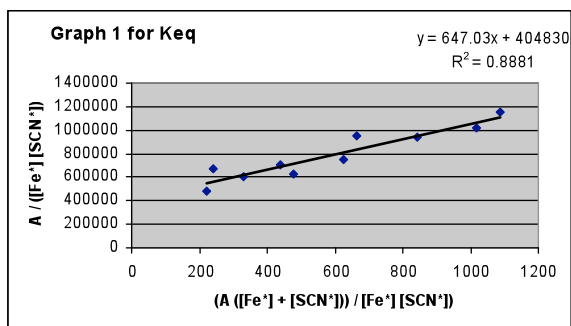
3. While doing the Beer's Law experiment, Bob and his lab partner used ~5ml of solution for each determination of the %T. Their classmate Suzy and her lab partner used ~10ml of solution for each determination. Explain briefly what differences if any would result from doubling the amount of solution.

There would be no difference in the transmittances observed because the volume in the cuvette does not matter. The pathlength and the concentration are what matter and they have not been changed.

4. While doing a Beer's Law experiment, Frank thought that he would run out of solution. Rather than go get more from the hood. He diluted his original solution to half its original concentration. How would this effect his absorbance readings compared to a student who had used the solution with the original concentration?

Since the absorbance is directly proportional to the concentration, reducing the concentration to half of the original concentration would result in the absorbance values being half of the original ones as well.

5. Trina's lab partner generated the following graphs for her and told her to finish the lab on her own.



Looking at the graphs,

a.) Determine the  $K_{eq}$  for this reaction.

$K_{eq}$  is equal to the slope of the trendline for the graph on the left, so  $K_{eq} = 647$ .

b.) Determine the concentration of an unknown solution when %T = 61.1.

(Assume the pathlength was 1.00cm.)

$$A = \log(100 / \%T) = \log(100 / 61.1) = 0.214$$

Then using the equation generated for the trendline for the graph on the right:  $y = 186.27x - 0.014$  where  $y =$  Absorbance and  $x =$  concentration, then:

$$x = (y + 0.014) / 186.27 = (0.214 + 0.014) / 186.27 = 0.228 / 186.27 = 0.00122 \text{ M}$$

**9. Atomic Spectra:** Using the Rydberg equation (where  $R = 3.29 \times 10^{15}$  Hz) and the speed of light ( $C = 2.998 \times 10^8$  m/s):

a. Calculate the expected frequencies in Hertz ( $s^{-1}$ ) of the radiation emitted by a hydrogen atom for the following electronic transitions.

$$\nu = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

b. Calculate the expected wavelengths in nanometers (nm) of the radiation emitted by a hydrogen atom for the same electronic transitions.

$$C = \lambda\nu$$

c. Label which wavelengths correspond to the Balmer series and which wavelengths correspond to the Lyman series.

Transitions	Frequency ( $s^{-1}$ )	Wavelength (nm)	Balmer / Lyman
$n_2 = 3$ & $n_1 = 1$	$2.92 \times 10^{15}$	$1.02 \times 10^{-7}$ m = 102 nm	Ultraviolet so Lyman
$n_2 = 2$ & $n_1 = 1$	$2.47 \times 10^{15}$	$1.21 \times 10^{-7}$ m = 121 nm	Ultraviolet so Lyman
$n_2 = 5$ & $n_1 = 2$	$6.91 \times 10^{15}$	$4.34 \times 10^{-7}$ m = 434 nm	Visible so Balmer
$n_2 = 4$ & $n_1 = 2$	$6.19 \times 10^{15}$	$4.84 \times 10^{-7}$ m = 484 nm	Visible so Balmer
$n_2 = 3$ & $n_1 = 2$	$4.57 \times 10^{15}$	$6.56 \times 10^{-7}$ m = 656 nm	Visible so Balmer

d. Why did the Hydrogen spectrum have the fewest lines?

The lines are created when the electrons make transitions from one level to another. Since Hydrogen has only one electron and since there are only so many allowable transitions, then the Hydrogen spectrum has the fewest lines.

e. For the Hydrogen spectra, why was the red line more intense (brighter) than the other lines?

The red line corresponds to the lowest energy transition in the visible spectrum (i.e.,  $3 \rightarrow 2$ ); so more electrons are making this transition because it takes the least amount of energy. Because more electrons are making the transition, it appears more intense.

**10. Flame Tests – What color flame is produced by each of the following elements?**

- copper – green
- iron – gold (orange/yellow)
- lithium – red
- potassium – purple
- sodium – yellow
- strontium – red

f. Why did we need the copper wire for the Beilstein reaction?

The copper in the wire reacts with the organohalides to produce a green flame.

**11. Gas Laws:** Using the ideal gas law calculate the volume of the system.

**Given:** pressure = 738 mmHg, mass = 0.725 grams,  $MW_{\text{butane}} = 58.000 \text{ g/mole}$ ,  $T = 20^\circ\text{C}$ ,  $R = 0.08206 \text{ Latm/molK}$

- a. What is the number of moles of butane?

$$0.725\text{g} (1 \text{ mole} / 58.000\text{g}) = 0.0125 \text{ moles of butane}$$

- b. What is the pressure in atm?

$$1 \text{ mmHg} = 1 \text{ torr} \qquad 760 \text{ torr} = 1 \text{ atm}$$

$$P = 738 \text{ torr} (1 \text{ atm} / 760 \text{ torr}) = 0.971 \text{ atm}$$

- c. What is the temperature in K?

$$T = (20 + 273.15) \text{ K} = 293.15 \text{ K}$$

- d. What is the volume of the system?

$$PV = nRT \qquad \text{so } V = nRT / P$$

$$V = (0.0125 \text{ mole}) (0.08206 \text{ L atm} / \text{mole K}) (293.15 \text{ K}) / (0.971 \text{ atm})$$

$$V = 0.309 \text{ L}$$

- e. What would the volume be at STP?

$$V = (0.0125 \text{ mole}) (0.08206 \text{ L atm} / \text{mole K}) (273.15 \text{ K}) / (1 \text{ atm})$$

$$V = 0.280 \text{ L}$$

or  $P_1V_1 / T_1 = P_2V_2 / T_2$

$$V_2 = P_1V_1T_2 / P_2T_1$$

$$V_2 = (0.971 \text{ atm}) (0.309 \text{ L}) (273.15 \text{ K}) / (1 \text{ atm}) (293.15 \text{ K})$$

$$V_2 = 0.280 \text{ L}$$

**12. Statistics:**

- a. For the following data set (2.10, 3.20, 3.50, 4.90, 4.30, 2.90) find the mean (average).

$$\bar{x} = (2.10 + 3.20 + 3.50 + 4.90 + 4.30 + 2.90) / 6$$

$$= 20.9 / 6$$

$$= 3.48$$

- b. For the average of the data set above, calculate the % Error if the expected answer was 3.500.

$$\% \text{Error} = [(3.500 - 3.483) / 3.500] \times 100$$

$$= 0.4857\%$$

- b. For this data set would you calculate the standard deviation or the standard deviation estimate?

Explain why.

You would use the standard deviation estimate because you had a small sample population.

**13. Dimensional Analysis:** Choose problems from sets 1, 2, 4 or 5 and work them.

**14 Scientific Notation & Significant Figures:**

- a. Choose problems from sets 1 & 2 and work them.  
b. Review problems from the midterm exam.

**\*\*Note:** Most of the questions on the final will be similar to those on review and on quizzes.



## 15. People – How did these people contribute to the experiments we did in Chem 2?

- a. Henri Becquerel (Nuclear) while studying fluorescence determined that some glowing rocks actually have particles coming off of them – the advent of radioactivity.
- b. Pierre and Marie Curie (Nuclear) were Nobel laureates who discovered elements Polonium and Radium which were more radioactive than uranium.
- c. Ernst Rutherford (Nuclear) developed an experiment where they shot  $\alpha$ -particles at gold foil. The particles scattered and it was realized the Plum Pudding model of the atom must not be accurate.
- d. Svante Arrhenius (Antacid) defined an acid as a substance that, when dissolved in water, increases the amount of hydronium ion over that present in pure water.
- e. Johannes Nicolaus Brønsted and Thomas Martin Lowry (Antacid) defines an acid as a substance that can donate a hydrogen ion.
- f. Gilbert N. Lewis (Antacid) defines an acid as any species that accepts electrons through coordination to its lone pairs.
- g. August Beer & Johann Heinrich Lambert (Equilibrium) showed that Absorbance is based the absorption (extinction) coefficient, the pathlength and the concentration of the solution.  $A=abc$
- h. Henry Louis Le Chatelier (Equilibrium) determined that when a reaction is shifted, the system will oppose any change in conditions to bring it back to equilibrium.
- i. Ibn Alhazen (Atomic Spectra) was the “Father of Optics” who studied light and determined many of its properties.
- j. Galileo Galilei (Atomic Spectra) turned his telescope to the heavens and discovered many astronomical phenomena which helped to prove that the earth travels around the sun.
- k. Sir Isaac Newton (Atomic Spectra) used prisms to prove white light was composed of colored light.
- l. Joseph von Fraunhofer (Atomic Spectra) studied the solar spectrum to try and improve glass lenses and discovered black lines in the solar spectrum.
- m. Bunsen & Kirchhoff (Atomic Spectra) figured out that the emission spectrum of elements matched the black solar lines seen by Fraunhofer.
- n. Johann Balmer (Atomic Spectra) mathematically described the series of lines in the spectrum for hydrogen.
- o. Max Planck (Atomic Spectra) quantized energy and described it mathematically with  $E = nh\nu$ .
- p. Neils Bohr (Atomic Spectra) developed the “planetary” model of the atom where electrons had fixed orbits.
- q. Robert Boyle (Gas Laws) “The Father of Modern Chemistry” was an Irish Chemist who found that at constant temperature  $P_1V_1 = P_2V_2$ .
- r. Jacques-Alexandre Charles (Gas Laws) was a French Chemist who found that at constant pressure  $V_1 / T_1 = V_2 / T_2$ .
- s. Amedeo Avogadro (Gas Laws) determined the number of atoms per mole; and found that at constant P&T, V is related to n.
- t. Joseph-Louis Gay-Lussac (Gas Laws) was a French Chemist who found that at constant volume  $P_1 / T_1 = P_2 / T_2$ .
- u. John Dalton (Gas Laws) showed that the partial pressure of each gas contributes to the total pressure.
- v. Johannes Diderik van der Waals (Gas Laws) revised the ideal gas law equation so that it can be used for real gases.

**Extra:** Albert Einstein was famous for his theories of relativity and  $E=mc^2$ ; but he also envisioned particles of light as photons.

- w. My TA’s name is...