Key for Sample Questions for Chem 002 Final Final WS13

1. **MSDS (the rest listed on review):**
    a. Proper attire – goggles, closed toe shoes, long pants or skirt or lab apron
    b. Acid Spill – neutralize with sodium bicarbonate
    c. Bunsen Burners – do not light if flammable reactants or products (e.g., H_2 gas) are present
    d. Phenolphthalein – has a laxative effect when ingested
    e. 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:**
    a. Determine if alpha or beta, then balance the following radioactive decay equations:

(alpha = ${}^{4}_{2}He$ / beta) $\quad {}^{222}_{86}Rn \rightarrow {}^{218}_{84}Po + {}^{4}_{2}He$

(alpha / beta = ${}^{0}_{-1}e$) $\quad {}^{234}_{90}Th \rightarrow {}^{234}_{91}Pa + {}^{0}_{-1}e + \textit{anti-}\upsilon$

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

$A = A_o e^{-kt}$ $\quad \ln A = -kt + \ln A_o$ $\quad \ln 2 = 0.693$ $\quad t_{½} = \ln 2 / k$ $\quad y = mx + b$ $\quad m = (y_2 - y_1) / (x_2 - x_1)$

and the data:

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

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

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

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

2. Determine the initial activity, A_o.

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

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

    $9.50 = -0.10 + \ln Ao$

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

3. 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 ΔH of this reaction.

Given: **C_p of solution (J/K) = 4.13 J/(g · K)*Volume of solution in mL** **(1 mL ≈ 1 g for aqueous soln)**

C_p of calorimeter (J/K) = 50

Q = (-total C_p* ΔT) **$\Delta T = T_f - T_i$**

ΔH = Q/n **n = # of moles reacted**

a. Determine the change in temperature for the system.

$\Delta T = Tf - Ti = 38.8\ ^{o}C - 24.6\ ^{o}C = 14.2\ ^{o}C$

Since we calculated ΔT, 14.2 °C = 14.2 K

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

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

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

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

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

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

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

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

(1.35 mole / L HCl) x (100 ml) x (1 L / 1000 ml) = 0.135 mole HCl

(1.76 mole / L NaOH) x (100 ml) x (1 L / 1000 ml) = 0.176 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 J/K) (14.2 K)]

Q = -12,439.2 joule

f. Determine the **change in enthalpy, ΔH**, for the reaction.

ΔH = (-12,439.2 joule) / (0.135 mole soln)

ΔH = -92,142 joule / mole

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

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

ΔH_{water} (J) $= C_{p,H2O}$ *(mass)*ΔT $\qquad$ ΔH_{ice} (J) $= C_{p,ice}$ *(mass)*ΔT $\qquad$ $H_{calorimeter}$ (J) $= C_{p,Cal}$ *ΔT

$C_{p,H2O} = 4.18$ J/(g$\cdot^o$C) $\qquad$ $C_{p,Cal} = 50$ J/^{o}C $\qquad$ $\Delta T = T_f - T_i$

a. Determine the ΔH_{water}.

ΔH_{water} (J) $= C_{p,H2O}$ *(mass)*ΔT
= 4.18 J/(g$\cdot^o$C) x (111.24 grams) x (15.8 ^{o}C – 23.2 ^{o}C)
= -3441 J

b. Determine the ΔH_{ice}. (***Hint** for* C_p – *The ice has melted.*)

Since $C_{p,ice} = C_{p,H2O}$

ΔH_{ice} (J) $= C_{p,H2O}$ *(mass)*ΔT
= 4.18 J/(g$\cdot^o$C) x (9.53 grams) x (15.8 ^{o}C – 0.0 ^{o}C)
= 629 J

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

$H_{calorimeter}$ (J) $= C_{p,Cal}$ *ΔT
= 50 J/^{o}C x (15.8 ^{o}C – 23.2 ^{o}C)
= -370 J

d. Determine the ΔH_{fus} for **one gram** of ice.
(***Hint***: For a calorimeter (i.e., closed systems) $\Delta H_{total} = 0$)

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

ΔH_{fus} = – ($\Delta H_{ice} + \Delta H_{water} + \Delta H_{calorimeter}$)
= – (629 J – 3441 J – 370 J)
= 3182 J

So ΔH_{fus} for **one gram** of ice

ΔH_{fus} = 3182 J / 9.53g
= 334 J/g

8. General Thermochemistry Concepts – a-d.) Circle appropriate answer. e-h.) Define**:**

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

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

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

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

e. This term means "the techniques that are used to measure enthalpy": **Calorimetry**

f. This term means "the energy needed to raise the temperature of an object 1^o C": **Heat Capacity**

g. This term means "the energy needed to raise the temperature of one gram of a substance 1^o C": **Specific Heat**

h. 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.

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

6. Antacids: You are given 1.12 M HCl and 1.48 M NaOH. The antacid you use contains 300 mg of $CaCO_3$ and 100 mg of $Al(OH)_3$. 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 mmole / ml HCl) x (35.0 ml HCl) = 39.2 mmole HCl

b. The number of **mmoles of NaOH** used to back titrate the acid.

(1.48 mmole / ml NaOH) x (21.8 ml NaOH) = 32.3 mmole NaOH

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

Excess HCl = mmole HCl – mmole NaOH = 39.2 -32.3 = 6.9 mmole HCl

d. Write the **balanced equations** for the neutralization of the antacid (Both $CaCO_3$ and $Al(OH)_3$).

$CaCO_3 + 2\ HCl \rightarrow CaCl_2 + H_2CO_3 \rightarrow CaCl_2 + H_2O + CO_2$

$Al(OH)_3 + 3\ HCl \rightarrow AlCl_3 + 3\ H_2O$

e. Using the **number of mg in the tablet**, calculate the mmoles of each component (Both $CaCO_3$ and $Al(OH)_3$).

300 mg $CaCO_3$ x (1 mmole / 100 mg $CaCO_3$) = 3.00 mmole $CaCO_3$

100 mg $Al(OH)_3$ x (1 mmole / 78 mg $Al(OH)_3$) = 1.30 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 mmole $CaCO_3$ x (2 mmole HCl / 1 mmole $CaCO_3$) = 6.00 mmole HCl

1.30 mmole $Al(OH)_3$ x (3 mmole HCl / 1 mmole $Al(OH)_3$) = 3.90 mmole HCl

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

6.00 + 3.90 mmole HCl = 9.9 mmole HCl

h. Calculate the **percent error** in order to compare the theoretical (g.) to the actual (c.). What are possible reasons this discrepancy could have occurred?

Actual is 6.9 mmole < Theoretical 9.9 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.

7. 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	-log (23.5/100) 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 (λ Max)

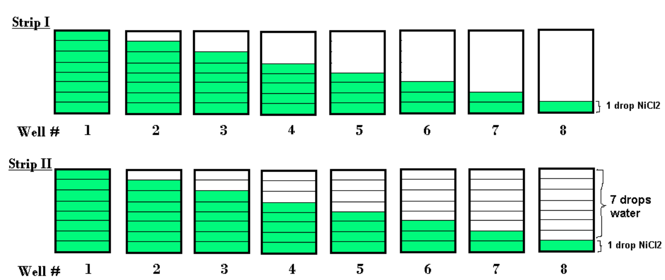
Blue Dye Max. Absorbance = 1.07 at 600 nm (λ Max)

c. Calculate the Absorbance Ratio of the Unknown/Standard at (λ Max).

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

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

8. 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.

a. What is changing in the first well strip – concentration or **pathlength**?

b. What is changing in the second well strip – **concentration** or pathlength?

c. Looking from the top how does the intensity compare for Strip 1 to Strip 2?

more intense – **the same** – less intense

d. Using $C_1V_1 = C_2V_2$, what is the approximate concentration in ppm for the unknown?

In Well #7 5.05ppm (2 drops) = C_2 (8 drops)

1.26(25) ppm = C_2

9. Equilibrium Constant & Beer's Law:

1. Why is it important that you avoid any loss of solutions when transferring and measuring solutions in an equilbrium 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 moles / y volume < x moles / (y-z) volume.
e.g., 1 mole / 0.4 L < 1 mole / 0.39 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.

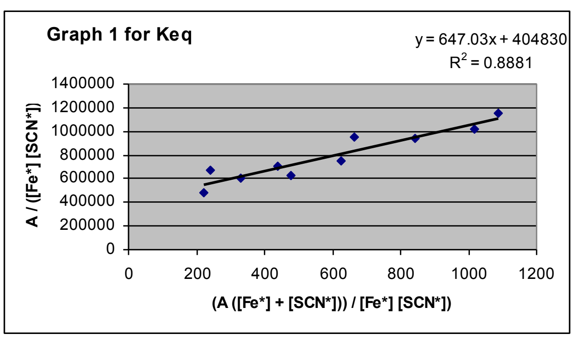
3. While doing the Beer's Law experiment, Bob and his lab partner used ~5ml of solution for each determination of the %T. Their classmate Suzy and her lab partner used ~10ml of solution for each determination. Explain briefly what differences if any would result from doubling the amount of solution.

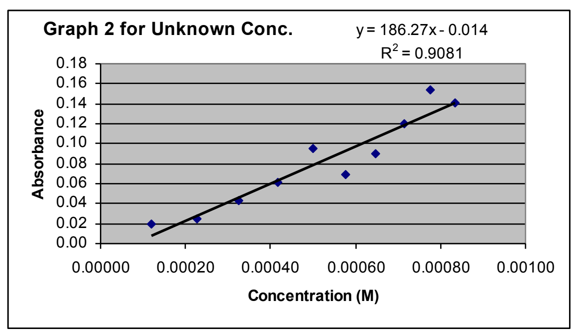
There would be no difference in the transmittances observed because the volume in the cuvette does not matter. The pathlength and the concetnration are what matter and they have not been changed.

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 Keq 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 M

10. Gas Laws: Using the ideal gas law, calculate the volume of the system.

Given: pressure = 738 mmHg, mass = 0.725 grams, MW_{butane} = 58.000 g/mole, T = 20°C, R = 0.08206 Latm/molK

a. What is the number of moles of butane?

0.725g (1 mole / 58.000g) = 0.0125 moles of butane

b. What is the pressure in atm?

1 mmHg = 1 torr   760 torr = 1 atm

P = 738 torr (1 atm / 760 torr) = 0.971 atm

c. What is the temperature in K?

T = (20 + 273.15) K = 293.15 K

d. What is the volume of the system?

PV = nRT   so V = nRT / P

V = (0.0125 mole) (0.08206 L atm / mole K) (293.15 K) / (0.971 atm)

V = 0.309 L

e. What would the volume be at STP?

V = (0.0125 mole) (0.08206 L atm / mole K) (273.15 K) / (1 atm)

V = 0.280 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 atm) (0.309 L) (273.15 K) / (1 atm) (293.15 K)

V_2 = 0.280 L

11. Statistics:

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

x_{bar} = (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.

%Error = [(3.500 – 3.483) / 3.500] x 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. (The standard deviation estimate represents a greater percentage of the population than the standard deviation. This is to compensate for the fact that a small sample population may not be truly representative of the whole population such that the true standard deviation then falls somewhere in the range of the estimate.)

12. Dimensional Analysis:

a. Choose problems from sets 1, 2, 4 or 5 and work them.

b. Dimensional analysis problems are generally incorporated within the other problems. For example:

1. Converting from mg to mmole in the antacid problem.
2. Converting from mmHg to torr or atm in the gas laws problem.
3. Converting from °C to K in the gas laws problem.

13. Scientific Notation & Significant Figures:

a. Choose problems form sets 1 & 2 and work them.

b. Review problems from the midterm exam. (*Monday midterm exam questions are shown below*).

_____ 22. What is the **numerical value** of 3.000×10^{1} ?

a. 0.3000 b. 3 c. 30 **d. 30.00**

_____ 23. **How many significant figures** are there in the number 0.09530?

a. 3 **b. 4** c. 5 d. 6

_____ 24. Which of the following numbers has <u>**3**</u> **significant figures**?

a. 0.1650 b. 2.030 **c. 0.0450** d. 318.0

_____ 25. Write 0.0025675 to <u>**4**</u> **significant figures**.

a. 0.0025 b. 0.003 **c. 2.568×10^{-3}** d. 2.5675×10^{3}

_____ 26. Using the **correct number of significant figures**, what is the answer to 1729.8 g + 2.29 g?

a. 1732.1 g b. 1732.09 g c. 1732 g d. 1730 g

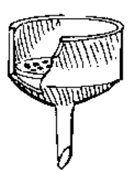
_____ 27. Using the **correct number of significant figures**, what is the answer when 6.5 is divided by 0.341 ?

a. 19.062 b. 19.06 c. 19.1 **d. 19**

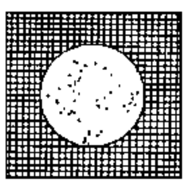
_____ 28. Find the **number of moles** in 100.0g of ammonia, NH_3.

a. 17.03 **b. 5.872** c. 0.1703 d. 3.535×10^{24}

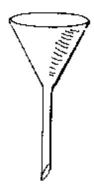
13. Glassware and equipment: Identify the equipment below.



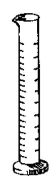
Buchner Funnel



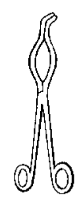
Wire Gauze



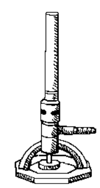
Short stemmed funnel



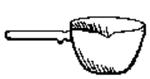
Graduated Cylinder



Crucible tongs



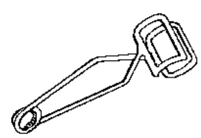
Bunsen Burner



Casserole

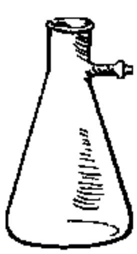
Evaporating dish

Watchglass



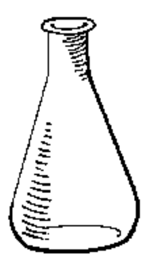
Test tube Clamp

Test tube brush

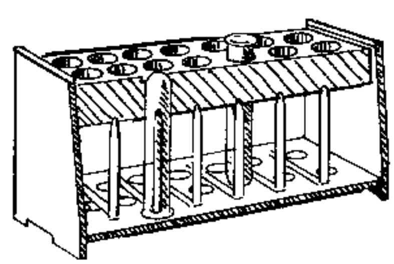


Vacuum flask

Beaker



Erlenmeyer flask



Test tube in test tube rack

14. People – How did these people contribute to the experiments we did in Chem 2?
(*All powerpoints are available at* http://web.mst.edu/~tbone)

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 α-particles at gold foil. The particles scattered and it was realized the Plum Pudding model of the atom must not be accurate.

d. Albert Einstein (Nuclear) was famous for his theories of relativity and $E=mc^2$; but he also envisioned particles of light as photons.

e. 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.

f. Johannes Nicolaus Brønsted and Thomas Martin Lowry (Antacid) defines an acid as a substance that can donate a hydrogen ion.

g. Gilbert N. Lewis (Antacid) defines an acid as any species that accepts electrons through coordination to its lone pairs.

h. August Beer & Johann Heinrich Lambert (Equilibrium) showed that Absorbance is is based the absorption (extinction) coefficient, the pathlength and the concentration of the solution. $A=abc$

i. 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.

j. Robert Boyle (Gas Laws) "The Father of Modern Chemistry" was an Irish Chemist who found that at constant temperature $P_1V_1 = P_2V_2$.

k. Jacques-Alexandre Charles (Gas Laws) was a French Chemist who found that at constant pressure $V_1 / T_1 = V_2 / T_2$.

l. Amedeo Avogadro (Gas Laws) determined the number of atoms per mole; and found that at constant P&T, V is related to n.

m. Joseph-Louis Gay-Lussac (Gas Laws) was a French Chemist who found that at constant volume $P_1 / T_1 = P_2 / T_2$.

n. John Dalton (Gas Laws) showed that the partial pressure of each gas contributes to the total pressure.

o. Johannes Diderik van der Waals (Gas Laws) revised the ideal gas law equation so that it can be used for real gases.

p. My TA's name is…

Ameya Natu	B1	F1
Andrew Powell	A1	E1
Anna Pfaff	C1	G1
Hooman Yaghoobnejad Asl	C2	G2
Jinyu Du	C3	G3
Jong-Sik Moon	B3	
Phalgun Lolur	A2	E2
Radheshyam Panta	A3	E3
Suraj Donthula	B2	F2

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