

Rogers

23 ✓  
19 ✓  
18 ✓  
12 ✓  
62

+ 1 OSD

63 students

## GENERAL CHEMISTRY II MIDTERM TEST

**INSTRUCTIONS: PLEASE READ THIS PAGE WHILE WAITING TO START YOUR EXAM.**

This test paper includes 4 pages (both sides) plus a periodic table; some potentially useful information is given on back of the periodic table. Check that your paper is complete before starting. You can remove the periodic table if you wish. Answer all questions inside the space provided. Calculators are permitted; cell phones and electronic dictionaries are not allowed. You have 75 min to complete the test. *I suggest you scan the whole test quickly before starting & do the 'easy' stuff 1<sup>st</sup>.*  
**GOOD LUCK!** Suggestion: spend 1 min / mark  $\Rightarrow$  25 min left to finish uncertain problems & check.

LAST NAME: marking scheme FIRST NAME: \_\_\_\_\_

STUDENT NUMBER: \_\_\_\_\_

90 min

Mark breakdown:

	Averages	Avg %
Page 2.	4.6 / 6	77%
Page 3.	4.0 / 6	67%
Page 4.	6.3 / 11	57%
Page 5.	8.8 / 11	80%
Page 6.	5.1 / 8	64%
Page 7.	4.2 / 9	47%

NOTE:  
✓ means 0.5 mark  
✓ means 1 mark  
X means wrong...

TOTAL: 32.9 / 50 (max. = 51)

PERCENT: 65.7 %

EARNED toward  
FINAL GRADE: 13.1 / 20

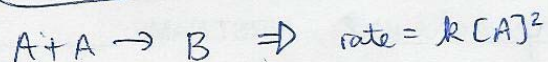
# 1. ( / 6 marks) Circle the words (IN CAPITALS) that correctly complete the following statements.  
[You do not need to explain your choices.]

- a) If a reaction that produces a gaseous product occurs in an open container, the expansion of the system causes the system to ( GAIN / LOSE ) energy in the form of work. Work is an example of a ( STATE FUNCTION / PATH FUNCTION ).

↳ lost energy by pushing back the surroundings (work = -PΔV)

- b) ( LONDON DISPERSION FORCES / DIPOLE-DIPOLE INTERACTIONS ) are intermolecular forces caused by transient ( EXCITATION / POLARIZATION ) of the electron clouds in neighbouring molecules.

- c) Imagine an elementary step that involves two molecules of the same compound. This process is a ( UNIMOLECULAR / BIMOLECULAR ) reaction, and it will display ( FIRST ORDER / SECOND ORDER ) kinetics.



- d) The normal boiling point of a substance is the temperature at which the substance's vapour pressure ( EQUALS / EXCEEDS ) the atmospheric pressure. For this reason, people living high in the mountains must cook their foods for a ( LONGER / SHORTER ) period of time than people living at sea level.

↳ lower atmospheric pressure, ∴ need less E input to reach bp. ∴ lower b.p.

↳ Since temperature is lower

- e) A catalyst ( DECREASES / INCREASES ) the rate constant of a reaction by decreasing the magnitude of the reaction's ( GIBB'S FREE ENERGY CHANGE / ACTIVATION ENERGY ).

$$k = Ae^{-E_a/RT} \quad \therefore \text{smaller } E_a \Rightarrow \text{larger } k$$

- f) To be spontaneous, a reaction that involves breaking a stronger covalent bond and forming a weaker covalent bond must be accompanied by a significant ( INCREASE / DECREASE ) in the randomness of the system.

$$\Delta H_{rxn}^{\circ} > 0 \quad \text{if new bond weaker than old bond}$$

$$\Delta G^{\circ} = \Delta H^{\circ} - T\Delta S^{\circ}$$

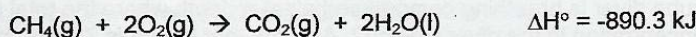
$$\ominus = \oplus - \underbrace{\quad}_{\oplus}$$

must have  $\Delta S^{\circ} > 0$  and  $T\Delta S^{\circ} > \Delta H^{\circ}$  to be spontaneous

6

- 0.25 for wrong SF (once per page maximum) (SF = significant figures)  
 - 0.25 for wrong unit analysis  
 - 0.25 rounding error

# 2. (6 marks) The hot water heaters in homes are often powered by burning natural gas. The combustion of methane (the principle component of natural gas), is described by the following equation:



[Note: potentially useful data is available on the formula page.]

- a) (3 marks) A typical shower uses 9.5 L of water per minute. Imagine you spend 8.5 minutes in the shower. What quantity of heat (in kJ) is required to warm the volume of water you use, if the water starts at 25°C and is heated in the hot water heater to 49°C?

$$\text{Volume of H}_2\text{O} = \frac{9.5 \text{ L}}{\text{min}} \times 8.5 \text{ min} = 80.75 \text{ L}$$

$$\text{mass of H}_2\text{O} = \left(\frac{1.000 \text{ g}}{1 \text{ mL}}\right) \times 80.75 \text{ L} \times \left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) = 80750 \text{ g}$$

$$q = m c \Delta T \quad \text{where } c = 4.184 \text{ J g}^{-1} \text{K}^{-1} \text{ and } 1 \text{ K} = 1^\circ \text{C}.$$

$$= (80750 \text{ g})(4.184 \text{ J g}^{-1} \text{C}^{-1})(49 - 25^\circ \text{C})$$

$$= 8108592 \text{ J}$$

$$\therefore q = 8100 \text{ kJ} \quad (2 \text{ SF})$$

NOTE: We are using the reaction only to provide the heat, not to MAKE the water! (#H<sub>2</sub>O is not relevant)

- b) (3 marks) If we assume the heat released by burning methane gas is transferred with 100% efficiency to the water, what volume of methane gas (at 25°C, 1.0 atm) will provide the amount of heat calculated in part (a)? [If you could not do part (a), use 13000 kJ...which is not right..to try (b).]

No heat lost  $\therefore$  need to find CH<sub>4</sub>(g) burned to yield 8110 kJ

$$\frac{1 \text{ mol CH}_4 \text{ burned}}{890.3 \text{ kJ released}} = \frac{x}{8110 \text{ kJ}}$$

$$x = 9.11 \text{ moles CH}_4 \text{ required}$$

Treat CH<sub>4</sub>(g) as an ideal gas to estimate volume:

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{(9.11 \text{ mol})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298.15 \text{ K})}{1.0 \text{ atm}}$$

$$= 223 \text{ L}$$

$$\therefore V \approx 220 \text{ L} \quad (2 \text{ SF})$$

Need 220 L of CH<sub>4</sub>(g) to yield enough heat.

- 0.25 SF errors

# 3. (11 marks) NOTE: This is not a Raoult's law calculation, but a partial pressure calculation. The Henry's law constant for nitrogen in blood serum is about  $8 \times 10^{-7} \text{ mol} \cdot \text{L}^{-1} \cdot \text{mmHg}^{-1}$ . Same math is not for same reasons!

a) (3 marks) If a scuba diver is breathing compressed air at a depth where the total pressure is 2.5 atm, what is the  $\text{N}_2$  concentration in the diver's blood? Assume that the air is 78%  $\text{N}_2$  by volume. done, but not for same reasons!

$P_{\text{N}_2} = 0.78 (2.5 \text{ atm}) = 1.95 \text{ atm}$  Since  $P$  of gas  $\propto n$  of gas and  $n \propto \text{Volume}$  from  $PV = nRT$ .

3

$C = kP$  Henry's law  
 $= (8 \times 10^{-7} \frac{\text{mol}}{\text{L} \cdot \text{mmHg}}) (1.95 \text{ atm}) (\frac{760 \text{ mmHg}}{1 \text{ atm}})$   
 $= 0.001186 \text{ mol/L}$

For comparison: at surface  
 $P_{\text{N}_2} = 0.78 \times 1 \text{ atm} = 0.78 \text{ atm} = 593 \text{ mmHg}$   
 $C = (8 \times 10^{-7} \frac{\text{mol}}{\text{L} \cdot \text{mmHg}}) (593 \text{ mmHg})$   
 $\therefore C = 5 \times 10^{-4} \text{ M}$   
typical  $\text{N}_2$  conc. in blood

b) (3 marks) If the diver returns to the surface too quickly, what will happen to the dissolved  $\text{N}_2$  in his blood, and why? Briefly explain the physical and/or chemical phenomena involved, not the biology. IF happens slowly, the extra  $\text{N}_2$  leaves slowly + does not form bubbles: not dangerous.

3

Because a gas's solubility is proportional to the partial pressure of that gas above the liquid's surface, and the partial pressure of  $\text{N}_2$  will drop suddenly back to  $\sim 0.78 \text{ atm}$ , the blood will become supersaturated with  $\text{N}_2$ . The  $\text{N}_2(g)$  will begin leaving the blood, likely so rapidly that bubbles will form in the diver's blood. This will continue until the  $[\text{N}_2]$  returns to the solubility limit at  $P_{\text{N}_2} = 0.78 \text{ atm}$ . (THE BUBBLES, OF COURSE, ARE NOT HEALTHY + CAUSE TISSUE DAMAGE, ETC...) + yes, they will also expand as they rise...

c) (5 marks) If the release of nitrogen gas from the blood is a first order process with a half-life of 9 minutes, how long will it take for 40% of the dissolved nitrogen to leave the diver's blood? [Note: in reality, this process is much more complicated than what I am suggesting here...]

For 1st order process,  $t_{1/2} = \frac{\ln 2}{k}$   
Thus:  $k = \frac{\ln 2}{t_{1/2}} = \frac{\ln 2}{9 \text{ min}} = 0.0770 \text{ min}^{-1}$

\*NOTE: If trying to use Henry's law constant  $k$  in place of the rate constant  $k$ , Max. 2.5.

5

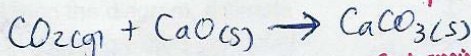
Using the 1st order integrated rate law:  
 $\ln [\text{N}_2]_t = \ln [\text{N}_2]_0 - kt$  where  $[\text{N}_2]_t = 0.60 [\text{N}_2]_0$  if 40% of  $\text{N}_2$  has left blood  
 $\ln(0.60 [\text{N}_2]_0) = \ln [\text{N}_2]_0 - kt$   
 $\ln(0.60 [\text{N}_2]_0) - \ln [\text{N}_2]_0 = -kt$   
log rules  $\Rightarrow \ln\left(\frac{0.6 [\text{N}_2]_0}{[\text{N}_2]_0}\right) = -kt$   
 $t = \frac{\ln(0.60)}{-k} = \frac{\ln 0.6}{-0.0770 \text{ min}^{-1}} = 6.634 \text{ min}$

$\therefore t \approx 7 \text{ min (ISF)}$

11

- # 4. ( / 11 marks) The carbon in the atmosphere (i.e.,  $\text{CO}_2$ ) is thermodynamically less stable than the carbon trapped in limestone (i.e.,  $\text{CaCO}_3$ ). In this problem, you will calculate thermodynamic quantities that will allow you to confirm or refute the validity of this statement.

- a) (1 mark) Carbon dioxide reacts with lime,  $\text{CaO(s)}$ , to produce  $\text{CaCO}_3(\text{s})$ . Write the balanced equation for this process.



Substance	$\Delta H_f^\circ$ (kJ/mol)	$S_f^\circ$ (J/mol·K)
$\text{CO}_2(\text{g})$	-393.51	213.74
$\text{CaO}(\text{s})$	-635.09	39.75
$\text{CaCO}_3(\text{s})$	-1206.92	92.9

- b) (3 marks) Calculate the standard enthalpy change for this reaction ( $\Delta H_{\text{rxn}}^\circ$ ). What effect would this reaction have on the entropy of its surroundings? Why?

$$\Delta H_{\text{rxn}}^\circ = \sum_n \Delta H_f^\circ \text{ products} - \sum_n \Delta H_f^\circ \text{ reactants}$$

$$= [(1 \text{ mol})(-1206.92 \text{ kJ}\cdot\text{mol}^{-1})] - [(1 \text{ mol})(-393.51 \text{ kJ}\cdot\text{mol}^{-1}) + (1 \text{ mol})(-635.09 \text{ kJ}\cdot\text{mol}^{-1})]$$

$$= -1206.92 \text{ kJ} - (-1028.60 \text{ kJ})$$

$$\therefore \Delta H_{\text{rxn}}^\circ = -178.32 \text{ kJ}$$

Heat is released to the surroundings  $\therefore$  disorder of surroundings  $\uparrow$   
 $\therefore$  increases entropy of surroundings

- c) (3 marks) Calculate the standard entropy change for this reaction ( $\Delta S_{\text{rxn}}^\circ$ ). Would the system become more organized or less organized when this reaction occurs?

$$\Delta S_{\text{rxn}}^\circ = [(1 \text{ mol})(92.9 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1})] - [(1 \text{ mol})(39.75 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}) + (1 \text{ mol})(213.74 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1})]$$

$$= 92.9 \text{ J}\cdot\text{K}^{-1} - (253.49 \text{ J}\cdot\text{K}^{-1})$$

$$\therefore \Delta S_{\text{rxn}}^\circ = -160.6 \text{ J}\cdot\text{K}^{-1}$$

The sign of  $\Delta S_{\text{rxn}}^\circ$  is negative,  $\therefore$  entropy of system has decreased  
 $\therefore$  system is less randomized now  
 $\therefore$  more organized after the reaction  
 (Note: matches logic since started with gas + now only have solid...)

- d) (4 marks) Calculate the standard Gibbs free energy change for this reaction at 298K. Does this confirm or refute the original statement that  $\text{CO}_2$  is thermodynamically unstable with respect to  $\text{CaCO}_3$ ? Why?

$$\Delta G_{\text{rxn}}^\circ = \Delta H_{\text{rxn}}^\circ - T \Delta S_{\text{rxn}}^\circ$$

$$= -178.32 \text{ kJ} - [(298 \text{ K})(-160.6 \text{ J}\cdot\text{K}^{-1}) \left( \frac{1 \text{ kJ}}{1000 \text{ J}} \right)]$$

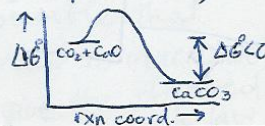
$$= -178.32 \text{ kJ} - (-47.86 \text{ kJ})$$

$$= -130.46 \text{ kJ}$$

$$\therefore \Delta G_{\text{rxn}}^\circ = -130.46 \text{ kJ}$$

The reaction is spontaneous in the forward direction. This implies that the reactants are less thermodynamically stable than the products, which CONFIRMS the statement made originally.

(Alternatively:  $\Delta G_f^\circ \text{CO}_2 = -393.51 \text{ kJ/mol}$  for 1 mol of C LESS STABLE  
 $\Delta G_f^\circ \text{CaCO}_3 = -1206.92 \text{ kJ/mol}$  for 1 mol of C MORE STABLE)



-0.25 SF errors

-1 for backwards Hess's law

(0.5 if state law correctly but enter #'s backwards!)

-0.5 no units

-0.25 wrong units

each species 0.25 (balanced...)

2 decimal places 1 decimal place

NOTE: mole units cancel out!

note: keep 2 decimal places as SF

mole units cancel out!

note: keep 1 decimal place as SF

sign errors (transcription) -0.25

Since has zero decimal places

but this should only have 2 SF, based on 298k...

because of subtraction rules for SF

- 0.25 SF errors  
 - 0.25 unit errors  
 - 0.5 no units

# 5. ( / 8 marks) Maple syrup is prepared by heating the sap from the maple tree to evaporate off much of the water. Heating is continued until the syrup's boiling point is about  $4.0^\circ\text{C}$  higher than the original sap. For this problem, we will approximate the syrup as being composed only of water and sugar ( $\text{C}_{12}\text{H}_{24}\text{O}_{12}$ ).

[Note: potentially useful data is available on the formula page.]

↳ a solution!

a) (2 marks) What is the molar concentration of sugar in maple syrup?

2

$$\Delta T_{bp} = K_b \cdot m \quad \text{where } m \text{ is the solute concentration in molality}$$

$$m = \frac{\Delta T_{bp}}{K_b}$$

$$= \frac{4.0^\circ\text{C}}{0.52^\circ\text{C}/m}$$

$$\therefore m = 7.69m$$

Because sugar is a molecular solute, the van't Hoff factor  $i=1$ , so the concentration of sugar is 7.7 molal. (2SF)

*extra 0.5 SF error*

[If you could not answer part (a), use a molal concentration of 5.0 m as a starting point for part (b).]

b) (6 marks) What mass of sugar is present in a 35 g portion of maple syrup? [Hint: what is the mass % composition of the syrup?]

1

$$\text{Mass \%} = \frac{\text{mass sugar}}{\text{mass sugar} + \text{mass H}_2\text{O}} \times 100\%$$

Link this to molality...

2

$$\text{molality} = \frac{7.69 \text{ moles sugar}}{1 \text{ kg solvent (H}_2\text{O)}}$$

note: kept 1 extra SF for calculations

3

Thus, can calculate mass of sugar + add it to mass H<sub>2</sub>O to find mass %.

Simply must calculate MM sugar  $\Rightarrow \text{MM}_{\text{C}_{12}\text{H}_{24}\text{O}_{12}} = 360.317 \text{ g/mol}$

$$\text{mass sugar in 1 kg H}_2\text{O} = (7.69 \text{ mol})(360.317 \text{ g/mol})$$

$$= 2770.8 \text{ g}$$

$$= 2770 \text{ g}$$

but remember, it is very sweet

$$\text{mass \% sugar} = \frac{2770 \text{ g sugar}}{(2770 \text{ g sugar} + 1000 \text{ g H}_2\text{O})} \times 100\%$$

definition worth 1.5

$$= 73.5\%$$

still keeping 1 extra SF since still doing more calculations

1

In a 35g portion of maple syrup, there would be

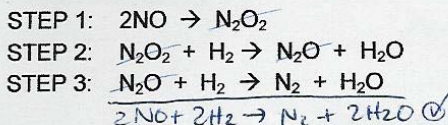
$$m = (0.735)(35\text{g}) = 25.72 \text{ g}$$

$$\therefore m = 26 \text{ g of sugar!}$$

(2SF)

8

- # 6. (9 marks) The reaction coordinate diagram below represents a proposed three-step mechanism for the reaction:  $2\text{NO}(\text{g}) + 2\text{H}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ .

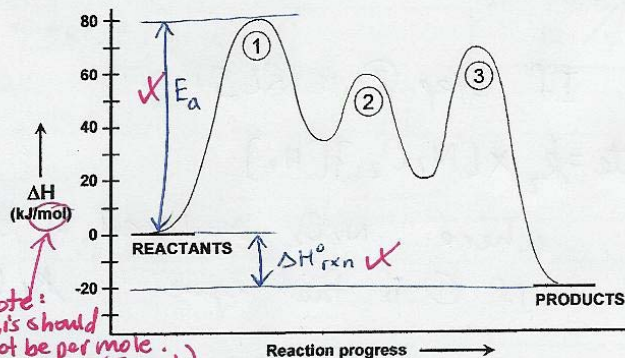


- a) (1 mark) Using the diagram, estimate  $\Delta H_{\text{rxn}}$  for the overall reaction, and label this quantity on the diagram.

$$\Delta H_{\text{rxn}} = \Delta H_{\text{prod}} - \Delta H_{\text{reactants}}$$

$$= (-20 - 0) \text{ kJ/mol}$$

$$= -20 \text{ kJ/mol}$$



- b) (2 marks) According to the diagram, which step is rate-limiting? Estimate the activation energy of the rate-limiting step, and label this quantity on the diagram.

Rate-limiting step = STEP ① since it has the largest activation energy.

$$E_a = \Delta H^\ddagger - \Delta H_{\text{reactants}}$$

$$= (+80 - 0) \text{ kJ/mol}$$

$$= 80 \text{ kJ/mol}$$

- c) (2 marks) Assuming the activation energy calculated in part (b) is correct, can you calculate the rate constant for the overall process at 298 K? Why or why not? [If you could not do (b), use  $E_a = 45 \text{ kJ/mol}$  for part (c).]

$$k = A e^{-E_a/RT}$$

We know  $E_a$ ,  $R$  &  $T$ , but not  $A$

$$= A e^{-\frac{80000 \text{ J}\cdot\text{mol}^{-1}}{8.314 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K} \cdot 298 \text{ K}}}$$

∴ No, we cannot calculate the rate constant.

∴  $k = A e^{-32}$  ...and we must now give up...

- d) (2 marks) What rate law is predicted for the reaction if it proceeds via this mechanism? Why?

Rate law for overall process = Rate law of rate-limiting step (RLS), (expressed in terms of reactants +/- products)

If RLS = step ①, then  $\text{rate} = k[\text{NO}]^2$  because step ① is a bimolecular step involving only NO.

(0.25 each...) must include!

- e) (2 marks) If the reaction is experimentally observed to be first order in  $\text{H}_2$ , would you describe the kinetics to be consistent with the proposed mechanism? Why or why not?

If rate experimental  $\propto [\text{H}_2]$ , then the rate-limiting step would need to involve  $\text{H}_2$ . Our proposed mechanism involves only NO in the RLS, so that would make our proposed mechanism inconsistent with the observed kinetics. It is possible that step ② or ③ might be the RLS (if so, our  $\Delta H$  diagram is inaccurate), but we would need to check to see if that would give the observed rate law.