

Rogers

GENERAL CHEMISTRY II MIDTERM TEST

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

This test paper includes 4 pages (both sides); 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 1st.*

GOOD LUCK! *Suggestion: spend 1 min / mark \Rightarrow 25 min left to finish uncertain problems & check.*

LAST NAME: marking scheme

FIRST NAME: _____

STUDENT NUMBER: _____

Mark breakdown:

	<u>Average</u>	<u>%</u>
Page 2.	6.3 / 10	63
Page 3.	6.9 / 11	63
Page 4.	4.2 / 10	42
Page 5.	5.4 / 10	54
Page 6.	4.6 / 10	46

*Too long!
Mark out of 44? ✓*

TOTAL: 27.5 / 50 (max. = 51)
BUT CALCULATED OUT OF 44, not 50.

PERCENT: 62.5 % ←

EARNED toward
FINAL GRADE: 12.5 / 20

1. (10 marks) True or False? Circle T or F to correctly describe each statement; you do not need explain your answers.

(T) / F A solution containing an ionic compound will have a lower vapour pressure than a solution containing the same concentration of a molecular compound.
 $\lambda > 1$ more particles of compound. $\lambda = 1$ fewer particles. $P_{vap} \downarrow$ as $[solute] \uparrow$ depends on $[particles]$

T / (F) Parts-per-million is a concentration unit based on molar quantities of solvent and solute.
mass mg solute / kg solution

(T) / F Increasing molecular mass often correlates with increasing intermolecular forces, because induced-dipole / induced-dipole interactions are influenced by the number and size of atoms present in the molecules. ✓

T / (F) When a gas-forming reaction occurs in a system, the system gains energy by doing work.
release gas $\Rightarrow \Delta V > 0$ System does work \therefore expends energy

(T) / F At higher elevations, where the atmospheric pressure is lower, it takes longer to cook foods in boiling water because the boiling point of water is below 100°C. ✓

(T) / F According to the second law of thermodynamics, a spontaneous change is accompanied by a decrease in the free energy of the system.
*spontaneous if $\Delta S_{univ} > 0$
 if $\Delta G < 0$, $\Delta S_{univ} > 0$ from definition of Gibbs free energy*

T / (F) Hydrophobic substances are not very soluble in water because the London dispersion forces between the hydrophobic molecules are much stronger than the dipole-induced-dipole interactions between the water and the hydrophobic molecules. ✓

T / (F) The vapour pressure above a mixture of two volatile liquids depends only on the vapour pressure of the liquid that is more volatile.
 $P_{vap} = X_1 P_1^0 + X_2 P_2^0$

(T) / F A reaction will be endothermic if it involves forming products with overall weaker bonds than the bonds in the reactants.
 $\Delta H > 0$ requires heat input.

T / (F) In general, if you study a reaction or process using a constant-volume ("bomb") calorimeter, the temperature change you measure will allow you to calculate the enthalpy change for your system.
if positive requires net Δ input

if const P: $\Delta E = \Delta H - P\Delta V$

otherwise: $\Delta E = q + w$

$\Rightarrow \Delta E = q_v$

$q_v = mC\Delta T$

$\Delta H = q_p$, not q_v
 by definition

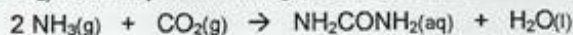
$-P\Delta V = 0$ if Volume is constant

BDE reactants (larger if stronger bonds) - BDE products (smaller if weaker bonds) > 0 gets released when form bonds.

✓ = 0.5 marks
 ✓ = 1 mark

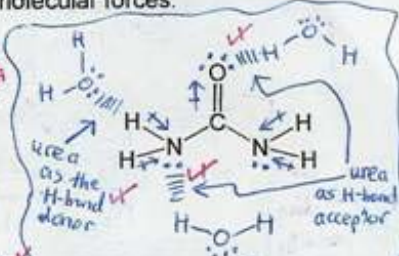
- 0.25 SF
 - 0.25 units (wrong or missing)
~~- 0.25 per mark (if not correct)~~

2. (/ 11 marks) Urea (NH_2CONH_2) is an important nitrogen-based fertilizer that is produced industrially by the reaction:



a) (3 marks) **MAXIMUM** Aqueous urea is the product of this reaction. (Explain why urea (shown below) is soluble in water, be sure to clearly identify (name AND sketch) the relevant intermolecular forces.

For a substance to be soluble, the interactions between solvent + solute molecules must be of similar strength as between solvent-solvent and solute-solute. Urea can undergo hydrogen bonding due to its $\text{N}-\text{H}$ bonds (N can accept; H can donate) and its $\text{C}=\text{O}$ bond (O can accept H-bonds). Water can therefore interact with urea just as favourably as it can with other water molecules, so dissolving urea is favourable. [Note that hydrogen-bonding is the strongest intermolecular force, so all others have been ignored.]



3 examples of hydrogen bonding interactions "H-bonds" with H_2O

b) (6 marks) Is the reaction of ammonia and carbon dioxide to produce aqueous urea thermodynamically favourable at 298 K? Show all necessary calculations. Need to show $\Delta G_{\text{rxn}}^\circ < 0$ for spontaneity.

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

Substance (at 298 K)	ΔH_f° (kJ/mol)	S_f° (J/mol·K)
$\text{NH}_3(\text{g})$	-45.90	192.77
$\text{CO}_2(\text{g})$	-393.51	213.74
$\text{NH}_2\text{CONH}_2(\text{aq})^*$	-317.1	172.8
$\text{H}_2\text{O}(\text{l})$	-285.83	69.95

(* Data adapted for the purpose of this exam.)

① $\Delta H_{\text{rxn}}^\circ = [(1\text{mol})(-317.1 \frac{\text{kJ}}{\text{mol}}) + (1\text{mol})(-285.83 \frac{\text{kJ}}{\text{mol}})] - [(2\text{mol})(-45.90 \frac{\text{kJ}}{\text{mol}}) + (1\text{mol})(-393.51 \frac{\text{kJ}}{\text{mol}})]$
 $= -602.92 - (-485.31) \text{ kJ}$
 $= -117.61 \text{ kJ}$ (with 1 extra SF)

Using Hess's law $\sum \Delta H_f^\circ \text{ prod} - \sum \Delta H_f^\circ \text{ reactants}$
 -0.5 if backwards (each time)
 ← net stabilization by forming products

② $\Delta S_{\text{rxn}}^\circ = [(1\text{mol})(172.8 \frac{\text{J}}{\text{mol}\cdot\text{K}}) + (1\text{mol})(69.95 \frac{\text{J}}{\text{mol}\cdot\text{K}})] - [(2\text{mol})(192.77 \frac{\text{J}}{\text{mol}\cdot\text{K}}) + (1\text{mol})(213.74 \frac{\text{J}}{\text{mol}\cdot\text{K}})]$
 $= 242.75 - 599.28 \text{ J/K}$
 $= -356.53 \text{ J/K}$ (with 1 extra SF)

③ $\Delta G_{\text{rxn}}^\circ = -117.61 \text{ kJ} - (298 \text{ K})(-356.53 \frac{\text{J}}{\text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}})$
 $= -11.4 \text{ kJ}$
 OR -11 kJ

equation ✓
 rounds to 106K (no decimal places significant now...)
 but still acceptable if treated 298K as "exact" here
 ⇒ Because the $\Delta G_{\text{rxn}}^\circ < 0$, the rxn is thermodynamically favourable.

c) (2 marks) Would this reaction give a higher yield of products at higher temperatures or at lower temperatures? Explain your choice.

Higher yield of products results when there is more driving force, i.e. when $\Delta G_{\text{rxn}}^\circ$ is a larger negative number. Because $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$, and because the reaction is enthalpically favourable ($\Delta H^\circ < 0$) but entropically unfavourable ($\Delta S^\circ < 0$), an increase in temperature will make the $-T\Delta S$ term a larger positive number. Thus, it counteracts the negative ΔH term and makes ΔG less negative (reduces driving force). So, to get a higher yield, we need a LOWER temperature, so that the energy requirement of organizing the system is small compared to the stabilization offered by the bonding.

-0.25 Significant figures error (at 18.82)
 -0.25 rounding error caused by not keeping extra SF during earlier calculations or by rounding molar masses, etc...

3. (2 / 10 marks) Ethylene glycol ($\text{HOCH}_2\text{CH}_2\text{OH}$) is commonly used as an antifreeze. Pure ethylene glycol is a liquid with a density of 1.11 g/mL . $\text{C}_2\text{H}_6\text{O}_2$

a) (6 marks) What volume (in L) of ethylene glycol would you add to a car radiator containing 6.50 L of water if the coldest winter temperature expected is -35°C ? [See helpful data on formula page...]

→ Aim to have solution freeze at -35°C (or lower), vs. 0°C for water
 i.e. $\Delta T_f = 35^\circ\text{C}$, a depression by this amount compared to water

2 $\Delta T_f = K_f \cdot m$
 $m = \frac{\Delta T_f}{K_f} = \frac{35}{1.86} = 18.82 \text{ mol} \cdot \text{kg}^{-1}$ (molality)

finding molality

we must convert molality into a practical quantity of solute for 6.50L of water... (note: this is the water volume, not the total volume)

2 Scaling for amount of H₂O present...
 Molality = $\frac{18.82 \text{ mol solute}}{1 \text{ kg of H}_2\text{O}}$, but we have $6.50 \text{ L H}_2\text{O} \times (1.00 \text{ g/mL})$
 $= 6.50 \text{ L} \times \frac{1.00 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$
 $= 6.50 \text{ kg H}_2\text{O}$
 $\therefore \frac{x \text{ mol}}{6.50 \text{ kg H}_2\text{O}} = \frac{18.82 \text{ solute}}{1 \text{ kg H}_2\text{O}} \Rightarrow x = 122.3 \text{ mol solute}$

2 So: solute mass required: $122.3 \text{ mol} \times 62.068 \text{ g/mol} = 7591 \text{ g}$ with $d = 1.11 \text{ g/mL}$
 solute volume: $\frac{7591 \text{ g}}{1.11 \text{ g/mL}} = 6839 \text{ mL}$
 Volume in Litres: $6839 \text{ mL} \approx 6.8 \text{ L}$ (2 SF)

b) (4 marks) Provide a detailed explanation of what causes a solution to have a lower freezing point than pure water. [Point form is ok.] TOTAL OF 4...

- At the freezing point: equal rates of freezing & melting (by definition)

• Effect of solute: ① interfere with freezing (slower freezing at any T)
 → since solvent molecules preferentially enter crystal
 → solute particles get in the way ∴ fewer solvent molecules can reach solid surface + enter solid phase i.e.: interactions between H₂O are disrupted

② facilitate melting (faster melting at any T)
 → crystal containing solute has defects in packing since solute molecules are different than neighbouring molecules in crystal
 → weaker overall crystal (weaker network of intermolecular forces)
 → requires less energy to overcome forces + melt

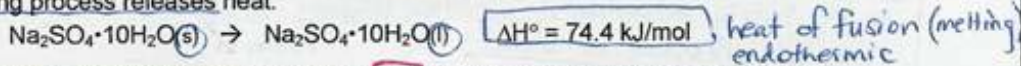
Overall: must go to a lower temperature than usual for melting to be slowed down sufficiently to balance with freezing rate ∴ freezing point is lower.

4 max.

10

-0.25 SF
-0.25 rounding

- # 4. (1/10 marks) Glauber's salt, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ (322.195 g/mol), melts at a convenient temperature of 32°C . Home-builders sometimes place plastic bags containing Glauber's salt in the ceilings of rooms to help regulate the temperature. During the day, the melting process absorbs heat from the room, and at night, the freezing process releases heat.



- a) (5 marks) How much heat must the air in a room lose for its temperature to drop by 8.2°C ?

Assume: (1) room's dimensions = $2.80 \text{ m} \times 10.6 \text{ m} \times 17.2 \text{ m}$. Note: $1 \text{ L} = 1 \text{ dm}^3$, $1 \text{ dm} = 0.1 \text{ m}$
(2) constant air density of 1.294 g/L ; air's specific heat capacity = $1.2 \text{ J/(g}^\circ\text{C)}$.

5

$$q = mC\Delta T \quad \checkmark \text{ where } \Delta T = -8.2^\circ\text{C} \text{ drop in temperature}$$

of air

$$C = 1.2 \frac{\text{J}}{\text{g}^\circ\text{C}}$$

$$= (660587 \text{ g}) \left(\frac{1.2 \text{ J}}{\text{g}^\circ\text{C}} \right) (-8.2^\circ\text{C})$$

$m_{\text{air}} = V \times d \leftarrow \text{must make units match here!}$

$$= -6500176 \text{ J}$$

$$q \approx -6500 \text{ kJ} \quad \checkmark \text{ should be -!}$$

\Rightarrow The air must lose 6500 kJ of heat.

(2SF)

$$= \left[\frac{(2.80 \times 10.6 \times 17.2) \text{ m}^3}{510.50 \text{ m}^3} \times \frac{1 \text{ L}}{(0.1 \text{ m})^3} \right] \times \left(\frac{1.294 \text{ g}}{1 \text{ L}} \right)$$
$$= (510.50 \text{ m}^3 \times \frac{1}{0.001 \text{ m}^3}) \times \frac{1.294 \text{ g}}{1 \text{ L}}$$

Volume = 510500 L

$$= 660587 \text{ g} \text{ of air in room.}$$

- b) (3 marks) What mass of Glauber's salt (in kg) is required to absorb the quantity of heat from (a)?

From $\Delta H_{\text{fus}}^\circ = 74.4 \frac{\text{kJ}}{\text{mol}}$, see that 1 mol of the salt would absorb 74.4 kJ

3

$$\text{So: } \frac{6500.18 \text{ kJ heat}}{x \text{ mol}} = \frac{74.4 \text{ kJ}}{1 \text{ mol}} \Rightarrow x = 87.4 \text{ mol} \quad (\text{extra SF})$$

$$\text{mass of Glauber's salt} = (87.4 \text{ mol}) (322.195 \text{ g} \cdot \text{mol}^{-1}) = 28149.5 \text{ g}$$
$$\approx 28 \text{ kg} \quad (2\text{SF})$$

\Rightarrow The ceiling would need 28 kg of salt in it to absorb sufficient heat to lower the temp by 8.2°C .

- c) (2 marks) In general, is the cooling of air an entropically favourable or entropically unfavourable process (according to the air)? Explain.

2

Air is a gas. (not actually important).

If a gas (or any substance) is cooled down, the average kinetic energy of its molecules/particles is decreased. This means the thermal motion is decreased, which causes a decrease in randomness. Thus, entropy decreases in the sample, which makes it an entropically unfavourable process.

10

- # 5. (1/6 marks) The osmotic pressure of human blood at body temperature (37°C) is 7.7 atm. If a solution containing equal concentrations of glucose ($C_6H_{12}O_6$) and sodium chloride (NaCl) is to be safely administered to a person intravenously, what must be the concentration (in molarity) of each substance?

To be safely introduced into the blood the solution must have the same osmotic pressure of blood; i.e. 7.7 atm ✓

This lets us calculate the total [solute particle] in the solution. ✓

$$\Pi V = nRT \Rightarrow \left(\frac{n}{V}\right) = \frac{\Pi}{RT} = \frac{7.7 \text{ atm}}{(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(37+273 \text{ K})} = 0.303 \frac{\text{mol}}{\text{L}} \quad \text{1 extra SF...}$$

So: what is in this solution? 1:1 $C_6H_{12}O_6$ and NaCl (same concentration) ✓
 i.e. molecular solute $C_6H_{12}O_6$ (i=1) AND ionic solute $\therefore Na^+$ cap & Cl^- cap (i=2) ✓

$$[\text{solute particles}] = 0.303 \text{ mol/L} = [C_6H_{12}O_6] + [Na^+] + [Cl^-]$$

because NaCl = 2

0.5 ✓ formed together from NaCl $\therefore [Na^+] = [Cl^-] = [NaCl]$

$$= [C_6H_{12}O_6] + [NaCl] + [NaCl]$$

$$\therefore 0.303 \text{ M} = [C_6H_{12}O_6] + 2[NaCl]$$

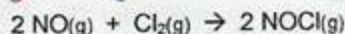
But, given that $[C_6H_{12}O_6] = [NaCl]$ as given in problem statement,

$$\therefore 0.303 \text{ M} = [C_6H_{12}O_6] + 2[C_6H_{12}O_6] = 3[C_6H_{12}O_6]$$

$$\text{Thus } \Rightarrow [C_6H_{12}O_6] = 0.101 \text{ M}$$

$$\therefore [NaCl] = 0.101 \text{ M} \quad \text{too } ([Na^+] = 0.101 \text{ M}, [Cl^-] = 0.101 \text{ M})$$

- # 6. (1/4 marks) The reaction shown at right was studied at -10°C:



The rate was determined by monitoring the Cl_2 concentration; the experimental data is shown in the table.

	[NO] ₀ (mol/L)	[Cl ₂] ₀ (mol/L)	Initial rate (mol/L·min)
①	0.10	0.10	0.18
②	0.10	0.20	0.36
③	0.20	0.20	1.45

- a) (2.5 marks) What is the rate law for this reaction?

Run ① → ②: $[Cl_2]_0 \times 2 \Rightarrow \text{rate} \times 2 \therefore$ 1st order in Cl_2

Run ② → ③: $[NO]_0 \times 2 \Rightarrow \text{rate} \times 2^2 \therefore$ 2nd order in NO

Thus: rate law is:

$$\text{rate} = k [Cl_2] [NO]^2$$

- b) (1.5 marks) What is the rate constant at this temperature? (You do not need to find the average.)

Choose any experimental data set: eg. run ①

$$\text{rate} = k [Cl_2] [NO]^2$$

$$\therefore k = \frac{\text{rate}}{[Cl_2][NO]^2} = \frac{0.18 \text{ mol L}^{-1} \text{ min}^{-1}}{(0.10 \text{ mol L}^{-1})(0.10 \text{ mol L}^{-1})^2}$$

$$\therefore k = 180 \text{ L}^2 \text{ mol}^{-2} \text{ min}^{-1} \quad \text{(Same as } \text{M}^{-2} \text{ min}^{-1})$$

no units? -0.5 Wrong units -0.25

10