Concordia University CHEM 206

Fall 2014, Dr. C. Rogers, Section 52

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FIRST NAME:	
STUDENT ID:	answer key

206 MT

F14

Chem 206 - GENERAL CHEMISTRY II

MIDTERM EXAMINATION

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

- This test includes 4 pages (both sides). Read over the whole test quickly before starting.
- Write answers in the space provided; fill in MCQ answers on the scan sheet in pencil.
- Calculators are permitted; cell phones and other electronic devices are not allowed.
- Potentially useful information is given on back of the (removable) periodic table.
- To earn full points, calculations must be annotated with explanatory comments.
- You have 70 min to complete the test.

- **GOOD LUCK!**

Mark breakdown:

avg. %		avg.	
62.1	Page 3.	8.7	/ 14
68.6	Page 4.	6.2	/ 9
51.8	Page 5.	5.2	/ ~~9~~ 10
50.7	Page 6.	4.6	/ 9

TOTAL: 24.6 / ~~40 (max = 41)~~ /41 max 42

But, % graded out of 40

Avg. PERCENT: 61.5 % (out of 40)

EARNED toward
FINAL GRADE: / 15

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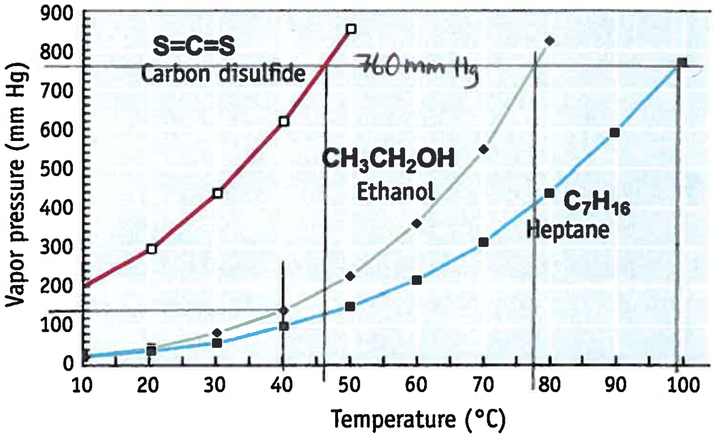
PART A: MCQs - MULTIPLE-CHOICE QUESTIONS (2 marks each)

Each question has one correct answer. Put answers in pencil on scan sheet; circle here as back-up.

Questions #1 - 4 involve the P-T graph shown.



#1. Which liquid is the least volatile?

B

- a) carbon disulfide
- [✓ b)] heptane
- c) ethanol
- d) all same

↳ lowest Pvap at all Ts, ∴ highest b.p. (hardest to evaporate)

#2. Which liquid has a normal boiling point of 46°C?

A

- [✓ a)] carbon disulfide
- b) heptane
- c) ethanol
- d) none

↳ Pvap = 760 mmHg 1 atm (technically 1 bar, but ~ same)

#3. For the liquid with the strongest intermolecular forces, what is its dominant intermolecular force?

A

- [✓ a)] London dispersion forces
- b) ion-dipole interactions
- c) dipole-dipole forces
- d) hydrogen bonding

↳ highest bp = heptane C_7H_{16} all C−C + C−H bonds (nonpolar) nonpolar molecule id - id forces only (dispersion)

#4. Imagine you dissolve a large amount of solute in a sample of liquid ethanol. At 40°C, what should the vapour pressure of ethanol be above this solution?

C

- ✗ a) cannot tell; need solute's identity
- b) more than 140 mm Hg
- [✓ c)] less than 140 mm Hg
- ✗ d) about 140 mm Hg

↳ pure ethanol Pvap 140 mmHg + solution will have LOWER Pvap than pure solvent

Questions #5 - 7 are based on the scenario below.

Mercury (II) chloride oxidizes oxalate anion via the reaction shown in the table. The method of initial rates was used (at 298 K) to collect the kinetics data shown.

$2\,HgCl_2 + C_2O_4^{2-} \rightarrow 2\,Cl^- + 2\,CO_2 + Hg_2Cl_2$

Initial reaction rates at 298 K

Run	$[HgCl_2]_o$ (M)	$[C_2O_4^{2-}]_o$ (M)	$Rate_o$ ($M{\cdot}min^{-1}$)
1	0.105	0.15	1.8×10^{-5}
2	0.105	0.30	7.2×10^{-5}
3	0.053	0.30	3.6×10^{-5}

÷2 ×2 ×2² ÷2¹

1st order in $HgCl_2$

2nd order in $C_2O_4^{2-}$

#5. What is the reaction's observed rate law?

D

- ✗ a) rate = $k\,[HgCl_2]^2[C_2O_4^{2-}]\,/\,[Cl^-]^2[CO_2]^2[Hg_2Cl_2]$
- ✗ b) rate = $k\,[HgCl_2][C_2O_4^{2-}]^{1}$ no.
- ✗ c) rate = $k\,[HgCl_2]^{2}[C_2O_4^{2-}]$
- [✓ d)] rate = $k\,[HgCl_2][C_2O_4^{2-}]^{2}$ yes

products are "never" in rate laws (very rarely)

#6. What is the value of the rate constant at 298 K?

B

- a) $1.1\times10^{-2}\ M^{-2}{\cdot}min^{-1}$
- [✓ b)] $7.6\times10^{-3}\ M^{-2}{\cdot}min^{-1}$
- c) $1.1\times10^{-3}\ M^{-2}{\cdot}min^{-1}$
- d) insufficient data provided

$$k = \frac{rate}{[HgCl_2][C_2O_4^{2-}]^2} = \frac{1.8\times10^{-5}\ M{\cdot}min^{-1}}{(0.105\,M)(0.15\,M)^2} = 7.6\times10^{-3}\ M^{-2}\,min^{-1}$$

#7. What is the rate of appearance of Hg_2Cl_2 when the rate of disappearance of $HgCl_2$ is 4.0×10^{-5} M/min?

C

- a) $8.0\times10^{-5}\ M{\cdot}min^{-1}$
- b) $4.0\times10^{-5}\ M{\cdot}min^{-1}$
- [✓ c)] $2.0\times10^{-5}\ M{\cdot}min^{-1}$
- d) $1.0\times10^{-5}\ M{\cdot}min^{-1}$

$$\frac{2\,HgCl_2}{4.0\times10^{-5}} \rightarrow \frac{1\,Hg_2Cl_2}{x}$$

$x = 2.0\times10^{-5}$ M/min

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PART B: CALCULATION & SHORT-ANSWER QUESTIONS

Show full calculations with units and brief comments to clarify the science behind the math.

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#8 **(9 marks)** Wet limestone is used to scrub (that is, chemically remove) sulfur dioxide gas from the exhaust gases of coal-burning power plants. The reactions shown below both occur, to form (i) hydrated calcium sulfite and (ii) hydrated calcium sulfate:

SO_2

(i) $CaCO_3(s) + SO_2(g) + \frac{1}{2}H_2O(\ell) \rightarrow CaSO_3{\bullet}\frac{1}{2}H_2O(s) + CO_2(g)$

(ii) $CaCO_3(s) + SO_2(g) + \frac{1}{2}H_2O(\ell) + \frac{1}{2}O_2(g) \rightarrow CaSO_4{\bullet}\frac{1}{2}H_2O(s) + CO_2(g)$

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

Thermodynamic data (at 298 K)		
Reaction	ΔH°_{rxn} (kJ·mol^{-1})	ΔS°_{rxn} (J·mol^{-1}·K^{-1})
(i)	-57.8	-39.8
(ii)	-320.9	-128.9

a) **(~~5~~4 marks)** For each reaction, use the data to calculate the Gibb's free energy of reaction at 298 K. Which reaction is more product favoured at this temperature? more ⊖ ΔG° ... ✓

(i) $\Delta G^\circ_{rxn} = \left(-57.8\ \frac{kJ}{mol}\right) - (298\ K)\left(-0.0398\ \frac{kJ}{mol \cdot K}\right)$ ✓ ✓

$= -57.8 + 11.86$ kJ/mol

$= -45.9$ ✓ kJ per mole product

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(ii) $\Delta G^\circ_{rxn} = \left(-320.9\ \frac{kJ}{mol}\right) - (298\ K)\left(-0.1289\ \frac{kJ}{mol \cdot K}\right)$ ✓ ✓

$= -320.9 + 38.41$ kJ/mol

$= -282.5$ ✓ kJ/mol product

Making $CaSO_4 \cdot \frac{1}{2}H_2O$ (ii) is more product-favoured at this T. ✓

−0.25 sf errors ⇒ ADDITION/SUBTRACTION RULES APPLY to $\Delta H - T\Delta S$ (keep lowest * decimal places)

~~−0.25 math error~~

b) **(1 mark)** In which reaction is the change in bond strength from reactants to products the largest? [illegible]

[illegible]

+ Could you predict this by looking at the reaction equation?

2

largest $|\Delta H^\circ_{rxn}|$ ✓ for rxn (ii) ✓

means largest difference in bond strengths.

Not easy to predict here ✓ → not even possible! ✓ because (ii) has more bonds broken (O_2 too...) BUT ends up with different ions in product

c) **(3 marks)** It is reasonable to say that exhaust gases from a power plant would be warmer than 298 K. Would you expect a higher yield of products or a lower yield of products from these reactions at higher temperatures? Briefly justify your choice. (If the two reactions behave differently, please explain why.)

3

Both reactions* are entropically unfavourable ✓ ($\Delta S^\circ_{rxn} < 0$) ✓. This means both reactions would yield LESS product ✓ at higher T, because the unfavourability of the entropy term will be amplified as T ↑ ✓:

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

✓

⊕ term will overwhelm favourable ⊖ ΔH term as T becomes larger + larger

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(* the effect will be most dramatic for (ii) because $|\Delta S|$ is larger) ✓

10 (typo)

#9. (~~9~~ marks) To be an effective radiator fluid (automobile coolant), a solution must remain liquid at the coldest expected winter temperatures (*e.g.*, -30°C). Consider some radiator fluid that was prepared by mixing 1.00 L of ethylene glycol ($HOCH_2CH_2OH$, d 1.114 g/mL) with 1.00 L of water at 25°C.

MM = 62.069 g/mol pure liquid

a) **(5 marks)** Calculate the freezing point of this radiator fluid, and comment on whether or not you think it would be an appropriate choice for winter use in Montreal. *Refer to data sheet for relevant properties of water.*

5

$\Delta T_{fp} = i\,m\,K_f$ ✓ (where $i = 1$ ✓ because ✓ eth. glycol is a molecular compound (non-electrolyte) ✓

and $m = \frac{n_{eth.glycol}}{\text{kg water}}$ ✓ $= \frac{(1.00\,L \times 1114\,\frac{g}{L})}{(62.069\,\frac{g}{mol})(1.00\,L_{H_2O} \times 1.000\,\frac{kg}{L})}$ ✓ ✓ ✓

$= \frac{17.948\text{ mol eth.glycol}}{1.00\text{ kg } H_2O}$ ✓

$= 17.948$ molal (keep 3SF later)

$K_f = 1.86\,°C\,\frac{kg\,H_2O}{mol\text{ solute}}$

not measured #, no effect on SF

$\Delta T_{fp} = (1)\left(17.948\,\frac{mol\ e.g.}{kg\,H_2O}\right)\left(1.86\,°C\,\frac{kg\,H_2O}{mol\,eg}\right)$ ✓

$= 33.38\,°C$ depression of f.p

∴ solution freezes at $-33.4\,°C$ (3SF) ✓

✓✓ Appropriate if T should remain above -30°C, because will still be liquid at -30°C + will flow through radiator properly.

b) **(2 marks)** Why does this solution not have the same freezing point as pure water? Briefly explain the molecular-level cause of this behaviour.

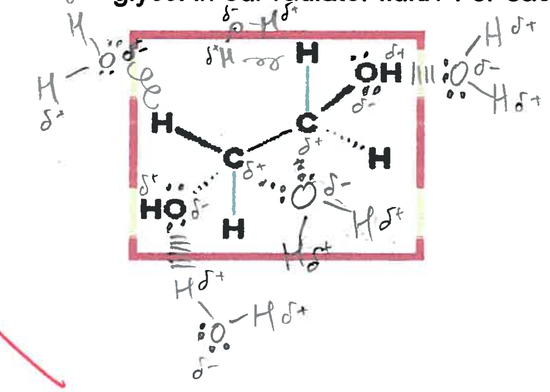
2

- Solute particles hinder freezing ✓ (solvent enters crystal preferentially) (✓)
  + facilitate melting ✓ (impure crystal ⇒ less KE to disrupt attractions) (✓)
- the more solute present, the greater the effect (✓) + the lower ✓ the T at which $rate_{freezing} = rate_{melting}$ (= f.p.)

* NOT same as saying that freezing slows down, since BOTH processes affected... (EQM...)

c) **(3 marks)** To the structure of ethylene glycol, add (i) lone pairs where appropriate and (ii) δ^+ δ^- labels to show bond polarities. What intermolecular forces exist between molecules of water and ethylene glycol in our radiator fluid? For each force, clearly indicate which parts of both molecules are involved.

3



✓ ① hydrogen bonding : see ≡ ✓ when O–H lined up with : on O

✓ ② dipole-dipole interactions : see ••• ✓ when at random angles (δ+ δ-)

✓ ③ London forces : see ~~ ✓ inducing dipoles when polar H_2O bumps into any nonpolarized atoms in molecule (CH's)

10

⊛ Do ~~not~~ not round off #s at each step. This causes a lot of error after a few steps. Keep 1-2 extra SF until the end.

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(−0.25) for premature rounding (−0.25) for SF errors (−0.25) units

PART B: CALCULATION & SHORT ANSWER QUESTIONS

/Show full calculations with units and brief comments to clarify the science behind the math.

(#10.) **(9 marks)** A 1.620 g of naphthalene ($C_{10}H_8$) is burned in the presence of excess oxygen in a bomb calorimeter with heat capacity of 7.721 kJ/°C. The temperature of the calorimeter increases by 8.44°C.

MM = 128.174 g/mol

a) **(4 marks)** Using the calorimetry data, calculate the energy released per mole of naphthalene burned.

1st: # mol naphthalene burned, $m = \frac{1.620\ g}{128.174\ g/mol} = 1.2639 \times 10^{-2}$ mol ✓ (4SF)

Bomb calorimeter = constant volume (✓); and all heat released by rxn is absorbed by calorimeter.

$\boxed{q_{v\,(burning)} = -q_{cal}}$ ✓

$= -\left(7.721\ \frac{kJ}{°C}\right)(+8.44°C)$ ✓ ✓ no mass given or needed; this is for the calibrated calorimeter as an object.

$= -65.165$ kJ from burning 1.620 g naphthalene. (3SF)

4

$\Delta E_{burning} = \frac{-65.165\ kJ}{1.2639 \times 10^{-2}\ mol}$ ✓ ✓ (✓) energy released per mole naphth.

$= -5155.9$ kJ/mol naphthalene

$\therefore \Delta E = -5.16 \times 10^{3}\ \frac{kJ}{mol}$ (3SF) $(\Rightarrow\ = \Delta H^{\circ}_{comb} + KE\uparrow$ due to trapped gas $(P\uparrow))$

b) **(2 marks)** Using the standard enthalpies shown in the table, calculate the standard molar enthalpy of combustion of naphthalene. Remember, the combustion of naphthalene occurs via the following balanced reaction equation: gas builds up: $\Delta n_{gases} = +2 \Rightarrow$ relevant to part (c)

$C_{10}H_8(s)$ + 12 $O_2(g)$ → 10 $CO_2(g)$ + 4 $H_2O(g)$

Thermodynamic data (at 298 K)		
Substance	**ΔH°$_f$ (kJ/mol)**	
$C_{10}H_8(s)$	+77.9	x 1 (R)
$O_2(g)$	by def'n: 0	x 12
$CO_2(g)$	-393.5	x 10 (P)
$H_2O(g)$	-241.8	x 4

ref. value, not measured "0", no effect on SF

SF: +/− rules keep least decimal places

2

$\Delta H^{\circ}_{comb.} = [10\ \cancel{mol}(-393.5\ \frac{kJ}{\cancel{mol}}) + 4(-241.8)] - [1(77.9) + 12(0)]$ ✓ ✓

$= -4902.2 - 77.9$ kJ ✓ ✓ (moles cancel)

$\therefore \Delta H^{\circ}_{comb.} = -4980.1$ kJ per mole naphthalene

$\Rightarrow$ = heat flow if P is constant (not here!)

c) **(3 marks)** Explain why your answers from parts (a) and (b) are different.

3

The calorimeter is a constant-V apparatus; that means it cannot expand or contract ✓ if gas is produced or consumed in the reaction. Here, 12 mol gas → 14 mol gas for the balanced combustion equation, so pressure will build up ✓ in the calorimeter and result in a temperature increase ✓. The rest of the T increase was due to the reaction being strongly exothermic ✓ due to bonding ✓ changes, as reflected in ΔH°_{comb}. ($\Delta E = \Delta H + \cancel{w}$ 0 here ✓ only if P is constant)

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