CHEMISTRY 102B/102C
Exam III
November 30, 2022
T. Hummel

NAME ____________________________

SIGNATURE ________________________

SECTION __________________________

FORM "A"

This exam is made up of an answer sheet, two cover sheets and 8 numbered pages. Below are instructions for coding the answer sheet. The last page of this exam contains some useful equations and constants, plus the periodic table.

On the answer sheet:

1. **Use #2 pencil. Erase cleanly**.
2. Print your **NAME** in the appropriate designated spaces, then blacken in the letter boxes below each printed letter, last name first, then your first name initial.
3. Fill in your university **ID** number under **STUDENT NUMBER**.
4. Under **SECTION** write the five-digit number that corresponds to your section designation, and then blacken in the corresponding number of boxes. **For 102B students,** the numbers are: BQ2 = 00012, BQ3 = 00013, BQ4 = 00014, BQ6 = 00016, BQ7 = 00017, BQ8 = 00018, BQA = 00021, BQB = 00022, BQC = 00023, BQD = 00024, BQG = 00027, BQH = 00028, BQI = 00029. **For 102C students**, the numbers are: CQ1 = 00031, CQ2 = 00032, CQ3 = 00033, CQ4 = 00034, CQ5 = 00035, CQ6 = 00036, CQ7 = 00037, CQ8 = 00038, CQ9 = 00039, CQA = 00041, CQB = 00042, CQC = 00043, CQE = 00045
5. Under **NETWORK ID** print your University Network ID beginning on the left-hand side with box #1, and then blacken in the corresponding letters, numbers and/or dashes under each character. Do not fill in a character for any unused boxes.
6. Under **TEST FORM** blacken the letter corresponding to the form designated on the upper left hand corner of the exam booklet.
7. Your TA's name should be printed for **INSTRUCTOR** and write your section number for **SECTION** in the lines provided.
8. **Sign** your name (do not print) on the line provided. Print your name underneath it.
9. **Mark** only one answer per question and do not use the answer sheet for scratch paper or make any stray marks on it. Erase cleanly if you wish to change an answer. The exam itself can be used for scratch paper.

Work carefully and efficiently. If your answer differs from one given in the last proper significant figure, mark that answer as correct and not the response "none of these". All questions are worth the same.

Solubility rules:

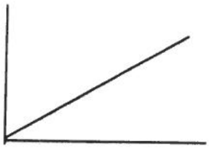
1. Most nitrate salts are soluble.
2. Most salts of alkali metals and ammonium cations are soluble.
3. Most chloride, bromide, and iodide salts are soluble.
   Exceptions: salts containing Ag^+, Pb^{2+}, and Hg_2^{2+} ions are insoluble.
4. Most sulfate salts are soluble.
   Exceptions: sulfates containing Ca^{2+}, Ba^{2+}, Pb^{2+}, and Hg_2^{2+} ions are insoluble.
5. Most hydroxide salts are insoluble.
   Exceptions: hydroxides containing alkali metals, Ba^{2+}, Sr^{2+}, and Ca^{2+} ions are soluble.
6. Most sulfide, carbonate, chromate, and phosphate salts are insoluble.
   Exceptions: salts of alkali metals and ammonium cations are soluble.

1. Initially, 1 mol of SO_3 is placed into a 2.0 L container at a constant temperature. Some SO_3 reacts to form SO_2 and O_2 by the following equilibrium reaction:

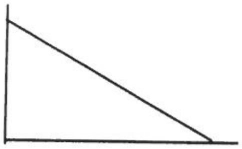
$$2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$$

Which of the following graphs most accurately represents the SO_3 concentration versus time plot for this equilibrium reaction? $[SO_3]$ is the y-axis, time is the x-axis.

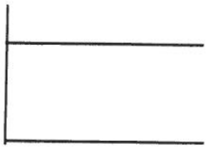
a)



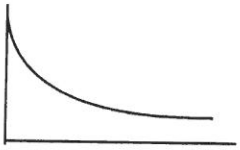
b)



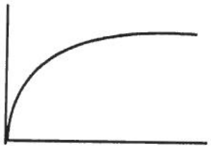
c)



d)



e)



2. Some A(g) is reacted by the following the reaction at a constant pressure of 1.0 atm:

$$A(g) \rightarrow 2\ B(g) \qquad \Delta H = -100.\ kJ$$

As a result of this reaction, the volume of the reaction container increases by 50.0 L and the internal energy changes by –30.0 kJ. How many moles of A were consumed in the reaction?

a) 1.25 mol b) 0.80 mol c) 0.35 mol d) 0.25 mol e) 0.20 mol

3. A system undergoes a process consisting of the following two steps:

Step 1: The system absorbs 40 J of heat while 60 J of work is done on it.
Step 2: The system absorbs 60 J of heat while performing 40 J of work.

Calculate ΔE for the overall process.

a) 200 J b) 120 J c) 0 J d) –120 J e) –200 J

Consider the following information for the next two questions. When 50.0 mL of 1.00 *M* $Ca(NO_3)_2$ is mixed with 50.0 mL of 1.00 *M* KF, a precipitate of $CaF_2(s)$ forms. K_{sp} for $CaF_2 = 4.0 \times 10^{-11}$.

4. Calculate the equilibrium concentration of Ca^{2+} in solution ($[Ca^{2+}] = ?$).

a) 1.00 *M* b) 0.75 *M* c) 0.50 *M* d) 0.25 *M* e) 0.050 *M*

5. Calculate the equilibrium concentration of F^- in solution ($[F^-] = ?$).

a) 4.5×10^{-6} *M* b) 1.3×10^{-5} *M* c) 8.9×10^{-6} *M*

d) 6.3×10^{-6} *M* e) 2.0×10^{-11} *M*

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6. At 1.00 atm and −78.0°C, carbon dioxide converts directly from the solid phase to the gaseous phase without going through the liquid phase; this is called sublimation. The sublimation process and energy change for $CO_2(s)$ at −78.0°C is:

$$CO_2(s) \rightarrow CO_2(g) \qquad \Delta H_{\text{sublimation}} = 25.3 \text{ kJ/mol}$$

Calculate the energy required to convert 88.0 g of $CO_2(s)$ at −78.0°C to $CO_2(g)$ at 25°C. The specific heat capacity of $CO_2(g)$ is 0.844 J/°C•g.

a) 3.99 kJ b) 3.99×10^3 kJ c) 58.3 kJ d) 7.70×10^3 kJ e) 54.5 kJ

7. Consider the following equilibrium constant vs. temperature data for some reaction:

K	Temp
2.54×10^4	109°C
5.04×10^2	225°C
6.33×10^1	303°C
2.25×10^{-1}	412°C
3.03×10^{-3}	539°C

Which of the following statements (a-c) is/are **true**?

a) The reaction is exothermic.
b) The reaction is endothermic.
c) As temperature increases, more products are produced in order to get to equilibrium.
d) Statements a and c are both true.
e) Statements b and c are both true.

8. Which of the following statements (a-d) is **true**?

a) If the reaction quotient Q is greater than K, the reaction must shift to the right to establish equilibrium.
b) The value of K can never equal the value of K_p.
c) At equilibrium, the rate of the forward reaction and the rate of the reverse reaction both equal zero.
d) A reaction contains mostly products at equilibrium when K << 1.
e) None of the above statements (a-d) are true.

9. Consider the following reaction at some constant temperature:

$$2\ SeO_3(g) \rightleftharpoons 2\ SeO_2(g) + O_2(g) \qquad K = 4.0 \times 10^{-12}$$

1.0 mol of SeO_3 is placed in a 1.0 L container, which then reacts to reach equilibrium according to the above reaction. At equilibrium, what is the concentration of $SeO_2(g)$?

a) $4.0 \times 10^{-12}\ M$
b) $2.0 \times 10^{-6}\ M$
c) $1.0 \times 10^{-6}\ M$
d) $1.0 \times 10^{-12}\ M$
e) $2.0 \times 10^{-4}\ M$

10. Consider the following solubilities of cerium iodate ($K_{sp} = 3.5 \times 10^{-10}$):

I. Solubility of $Ce(IO_3)_3(s)$ in water
II. Solubility of $Ce(IO_3)_3(s)$ in 0.10 *M* $Ce(NO_3)_3(aq)$
III. Solubility of $Ce(IO_3)_3(s)$ in 0.10 *M* $KIO_3(aq)$

Which of the above solubilities (I-III) is **smallest**?

a) I
b) II
c) III
d) All of the solublities (I-III) are the same.

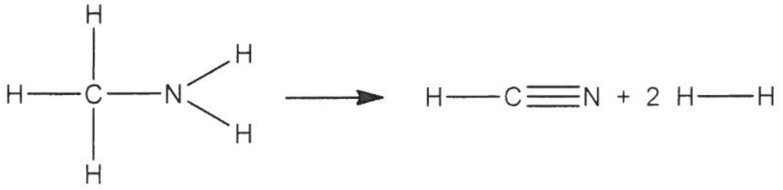
11. Consider the following data:

Specific heat capacity of ice = 2.03 J/°C•g
Specific heat capacity of water = 4.18 J/°C•g
Specific heat capacity of steam = 2.02 J/°C•g
ΔH_{fusion} = 6.02 kJ/mol
$\Delta H_{vaporization}$ = 40.7 kJ/mol

One mole of ice (18.0 g) at −5.0°C is added to one mole of water at 90.0°C. Calculate the final temperature of the mixture assuming no heat loss to the surroundings.

a) 3.8°C
b) 5.9°C
c) 12.6°C
d) 31.4°C
e) 68.5°C

12. Hydrogen cyanide can also be produced by the reaction:



$\Delta H° = 158$ kJ

Given the following bond energies, calculate the bond energy of the H–H bond.

	bond energy (kJ/mol)
H–C	413
H–N	391
C–N	305
C≡N	891

a) 316 kJ b) 207 kJ c) 432 kJ d) 158 kJ e) 734 kJ

13. The solubility of lead phosphate in 1.00 L of 0.010 *M* $Pb(NO_3)_2$ is 5.0×10^{-25} mol/L. Calculate the K_{sp} value for $Pb_3(PO_4)_2(s)$.

a) 1.3×10^{-48} b) 1.3×10^{-25} c) 1.0×10^{-48}

d) 1.1×10^{-268} e) 1.0×10^{-54}

14. Hydrogen cyanide can be produced by the following balanced reaction:

$$2\,NH_3(g) + 3\,O_2(g) + 2\,CH_4(g) \rightarrow 2\,HCN(g) + 6\,H_2O(g) \qquad \Delta H° = ?$$

Given the following standard enthalpies of formation:

	$\Delta H_f°$ (kJ/mol)
$NH_3(g)$	−46
$CH_4(g)$	−75
$HCN(g)$	135
$H_2O(g)$	−242

calculate ΔH° for this reaction.

a) 505 kJ b) −505 kJ c) −940. kJ d) −1424 kJ e) 940. kJ

15. Consider the following five solutions:

I. A solution which is 1.0×10^{-8} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.
II. A solution which is 1.0×10^{-7} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.
III. A solution which is 1.0×10^{-6} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.
IV. A solution which is 1.0×10^{-4} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.
V. A solution which is 1.0×10^{-3} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.

In how many of these five solutions will a precipitate of $Ag_2CrO_4(s)$ form? K_{sp} for $Ag_2CrO_4 = 9.0 \times 10^{-12}$.

a) 1
b) 2
c) 3
d) 4
e) 5; a precipitate will form in all five of the solutions (I-V).

16. Consider the following reaction for which K = 4.0 at some temperature.

$$2\,NF_3(g) \rightleftharpoons N_2(g) + 3\,F_2(g) \qquad K = 4.0$$

Which of the following sets of concentrations does **not** represent an equilibrium mixture for this reaction at this temperature?

a) $[NF_3] = 1.0\ M$, $[N_2] = 4.0\ M$, $[F_2] = 1.0\ M$
b) $[NF_3] = 2.0\ M$, $[N_2] = 2.0\ M$, $[F_2] = 2.0\ M$
c) $[NF_3] = 1.0\ M$, $[N_2] = 0.50\ M$, $[F_2] = 2.0\ M$
d) $[NF_3] = 3.0\ M$, $[N_2] = 4.5\ M$, $[F_2] = 2.0\ M$
e) $[NF_3] = 2.0\ M$, $[N_2] = 8.0\ M$, $[F_2] = 1.0\ M$

17. Consider the generic reaction $A(g) \rightarrow B(g)$. If this reaction is run in a constant volume container, then the heat released or gained when 1 mol of A(g) is reacted is equal to:

a) ΔH
b) ΔE
c) ΔV
d) ΔT
e) w

18. Which of the following processes (a-c) is/are endothermic?

a) Natural gas is burned in a furnace.
b) When NH_4NO_3 dissolves in water, the solution gets colder.
c) Two oxygen atoms react to form the O_2 molecule.
d) Both processes b and c are endothermic.
e) All of the processes (a-c) are endothermic

19. Which of the following diatomic molecules has the largest bond energy?

a) Br_2
b) F_2
c) N_2
d) Cl_2
e) I_2

20. Using the following reactions and ΔH values, calculate ΔH_{rxn} for:

$$ClF(g) + F_2(g) \rightarrow ClF_3(g) \qquad \Delta H_{rxn} = ?$$

$2\ ClF(g) + O_2(g) \rightarrow Cl_2O(g) + F_2O(g)$ $\Delta H = 167.4$ kJ/mol

$2\ ClF_3(g) + 2\ O_2(g) \rightarrow Cl_2O(g) + 3\ F_2O(g)$ $\Delta H = 341.4$ kJ/mol

$2\ F_2(g) + O_2(g) \rightarrow 2\ F_2O(g)$ $\Delta H = -43.4$ kJ/mol

a) −217.5 kJ/mol b) −130.2 kJ/mol c) 217.5 kJ.mol

d) −108.7 kJ/mol e) 108.7 kJ/mol

21. Consider the following reaction at some temperature:

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

An equilibrium mixture contains 2.0 mol $CaCO_3(s)$, 4.0 mol $CaO(s)$, and 0.50 mol $CO_2(g)$ in a 2.0 L container. If at this same temperature another equilibrium mixture contains 6.0 mol $CaCO_3(s)$ and 4.0 mol $CaO(s)$ in a 2.0 L container, determine the equilibrium concentration of $CO_2(g)$ present in this mixture.

a) 0.25 *M* b) 0.33 *M* c) 0.50 *M* d) 0.75 *M* e) 1.50 *M*

22. The gas arsine, AsH_3, decomposes by the following reaction:

$$2\ AsH_3(g) \rightleftharpoons 2\ As(s) + 3\ H_2(g)$$

In an experiment, pure $AsH_3(g)$ of unknown initial concentration is placed into a 1.0 L container. After equilibrium is reached, 3.0 moles of $H_2(g)$ is produced and 6.0 moles of $AsH_3(g)$ is also present. Determine the initial concentration of $AsH_3(g)$ before the reaction took place ($[AsH_3]_{initial} = ?$).

a) 9.0 *M* b) 8.0 *M* c) 3.0 *M* d) 10.5 *M* e) 15.0 *M*

23. Consider the following reaction at some constant temperature:

$$I_2(g) + H_2(g) \rightleftharpoons 2\,HI(g) \qquad K = 0.75$$

Initially, 2.0 mol of I_2, 2.0 mol of H_2 and 2.0 mol of HI are mixed together in a 1.0 L rigid container and are allowed to react to reach equilibrium. Which of the following statements is **true** once this reaction reaches equilibrium?

a) At equilibrium, $[I_2] > [HI]$.
b) At equilibrium, $[I_2] = 2.0\ M$.
c) At equilibrium, $[I_2] > [H_2]$.
d) At equilibrium, $[H_2] < 2.0\ M$.
e) At equilibrium, $[HI] = 2.5\ M$.

24. How many of the following four statements (I-IV) about enthalpy is/are **true**?

I. Enthalpy is a state function.
II. Enthalpy is an extensive property.
III. Hess's law can be applied to determine the enthalpy change for a reaction.
IV. The standard enthalpy of formation for an element in its standard state is zero.

a) 0 (None are true.)
b) 1
c) 2
d) 3
e) 4 [All of the statements (I-IV) are true.]

25. Consider the following salts:

$AgOH(s)$, $K_{sp} = 1.0 \times 10^{-8}$
$Cu(OH)_2(s)$, $K_{sp} = 4.0 \times 10^{-18}$
$Cr(OH)_3(s)$, $K_{sp} = 2.7 \times 10^{-19}$

Place these salts in order of **increasing** molar solubility (from least soluble to most soluble).

a) $AgOH < Cu(OH)_2 < Cr(OH)_3$
b) $Cu(OH)_2 < AgOH < Cr(OH)_3$
c) $Cu(OH)_2 < Cr(OH)_3 < AgOH$
d) $Cr(OH)_3 < Cu(OH)_2 < AgOH$
e) $Cr(OH)_3 < AgOH < Cu(OH)_2$

26. Consider a theoretical salt which is composed of M^{4+} and X^- ions. If the solubility of the MX_4 salt is 1.00×10^{-2} mol/L, calculate the K_{sp} value for $MX_4(s)$.

a) 1.08×10^{-10}
b) 2.56×10^{-8}
c) 2.70×10^{-7}
d) 4.00×10^{-6}
e) 1.00×10^{-8}

27. A 10.0 g-sample of a salt is dissolved in 100.0 g of water at an initial temperature of 25.0°C. After the salt dissolves, the final temperature of the solution increases to 37.2°C. Assume the heat capacity of the solution is 4.184 J/°C•g and assume no heat is lost to the calorimeter. Given the following enthalpies of dissolution and molar masses, which salt (a-e) was dissolved?

	Salt	ΔH (kJ/mol)	Molar Mass (g/mol)
a)	NaCl	3.00	58.45
b)	NaI	−7.53	149.9
c)	KBr	19.9	119.0
d)	LiBr	−48.8	86.84
e)	KCl	17.2	74.55

28. Consider the following reaction at equilibrium:

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g) \qquad \Delta H = -92\ \text{kJ}$$

Which of the following changes will shift the equilibrium to the right (to products)?

1. increasing the temperature
2. decreasing the temperature
3. increasing the volume of the reaction container
4. decreasing the volume of the reaction container
5. removing some NH_3
6. adding some NH_3

a) 1, 4, 6   b) 2, 3, 5   c) 1, 3, 6   d) 2, 4, 5   e) 1, 3, 5

29. The standard enthalpy of formation for $H_2O(l)$ is −285.8 kJ/mol. Calculate the change in internal energy for the following process at 298 K and 1 atm:

$$H_2O(l) \rightarrow H_2(g) + 1/2\ O_2(g) \qquad \Delta E° = ?$$

a) 282.1 kJ   b) 289.5 kJ   c) 285.8 kJ   d) −282.1 kJ   e) −289.5 kJ

30. My answers for this Chemistry 102 exam should be graded with the answer sheet associated with:

a) Form A   b) Form B   c) Form C   d) Form D   e) Form E

USEFUL CONSTANTS/EQUATIONS

$K = °C + 273$

$PV = nRT$

$R = 0.08206$ L atm/K mol

1 L = 1000 mL

Avogadro's number, $N = 6.022 \times 10^{23}$

kinetic energy = $(1/2)\, mv^2$

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

$K_p = K(RT)^{\Delta n}$

$1\ J = 1\ kg\ m^2/sec^2$

$R = 8.3145$ J/K mol

$\Delta E = q + w$

101.3 J = 1 L • atm

$\Delta H = \Delta E + \Delta(PV)$

$w = -P\Delta V = -RT\Delta n$ (at constant P and T)

$\Delta E_{univ} = \Delta E_{surr} + \Delta E_{sys}$

$$\Delta H^o_{rxn} = \Sigma \Delta H^o_{f,\ products} - \Sigma \Delta H^o_{f,\ reactants}$$

$q = s \times mass \times \Delta T$, s = specific heat capacity

PERIODIC TABLE OF THE ELEMENTS

26
Fe
55.85

←Atomic number

←Atomic mass

1 1A	2 2A	3	4	5	6	7	8	9	10	11	12	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.70	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57 La* 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po 209	85 At 210	86 Rn 222
87 Fr 223	88 Ra 226	89 Ac† 227	104 Rf 261	105 Db 262	106 Sg 266	107 Bh 262	108 Hs 265	109 Mt 266	110 Ds 271	111	112						

*Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm 145	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
†Actinides	90 Th 232.0	91 Pa 231	92 U 238	93 Np 244	94 Pu 242	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 252	100 Fm 257	101 Md 258	102 No 259	103 Lr 260