2005 U. S. NATIONAL CHEMISTRY OLYMPIAD

LOCAL SECTION EXAM

Prepared by the American Chemical Society Olympiad Examinations Task Force

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DIRECTIONS TO THE EXAMINER

This test is designed to be taken with an answer sheet on which the student records his or her responses. All answers are to be marked on that sheet, not written in the booklet. Each student should be provided with an answer sheet and scratch paper, both of which must be turned in with the test booklet at the end of the examination. Local Sections may use an answer sheet of their own choice.

The full examination consists of 60 multiple-choice questions representing a fairly wide range of difficulty. Students should be permitted to use non-programmable calculators. A periodic table and other useful information are provided on page two of this exam booklet for student reference.

Suggested Time: 60 questions—110 minutes

DIRECTIONS TO THE EXAMINEE

DO NOT TURN THE PAGE UNTIL DIRECTED TO DO SO.

This is a multiple-choice examination with four choices for each question. There is only *one* correct or best answer to each question. When you select your choice, blacken the corresponding space on the answer sheet with your pencil. Make a heavy full mark, but no stray marks. If you decide to change your answer, be certain to erase your original answer completely.

Not valid for use as an ACS Olympiad Local Section Exam after March 28, 2005. STOCK CODE OL05

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ABBREVIATIONS AND SYMBOLS					
ampere	A	Faraday constant	*F*	molal	*m*
atmosphere	atm	formula molar mass	*M*	molar	M
atomic mass unit	u	free energy	*G*	molar mass	*M*
atomic molar mass	*A*	frequency	ν	mole	mol
Avogadro constant	N_A	gas constant	*R*	Planck's constant	*h*
Celsius temperature	°C	gram	g	pressure	*P*
centi– prefix	c	heat capacity	C_p	rate constant	*k*
coulomb	C	hour	h	retention factor	R_f
electromotive force	*E*	joule	J	second	s
energy of activation	E_a	kelvin	K	temperature, K	*T*
enthalpy	*H*	kilo– prefix	k	time	*t*
entropy	*S*	liter	L	volt	V
equilibrium constant	*K*	milli– prefix	m		

CONS
$R = 8.314$ J·mo
$R = 0.0821$ L·atm·mo
$1\ F = 96{,}500$ C·mol^{-1}
$1\ F = 96{,}500$ J·V^{-1}·mol^{-1}
$N_A = 6.022 \times 10^{23}$ mol^{-1}
$h = 6.626 \times 10^{-34}$ J·s
$c = 2.998 \times 10^{8}$ m·s^{-1}
0 °C = 273.15 K
1 atm = 760 mmHg

EQUATIONS

$$E = E^\circ - \frac{RT}{nF}\ln Q \qquad \ln K = \left(\frac{-\Delta H}{R}\right)\left(\frac{1}{T}\right) + \text{constant} \qquad \ln\left(\frac{k_2}{k_1}\right) = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

PERIODIC TABLE OF THE ELEMENTS

1 1A																	18 8A
1 **H** 1.008	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	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	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	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** 47.88	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.61	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** 183.8	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** (263)	107 **Bh** (262)	108 **Hs** (265)	109 **Mt** (266)	110 (269)	111 (272)	112 (277)		114 (2??)				

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
90 **Th** 232.0	91 **Pa** 231.0	92 **U** 238.0	93 **Np** (237)	94 **Pu** (244)	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** (262)

DIRECTIONS

- When you have selected your answer to each question, blacken the corresponding space on the answer sheet using pencil. Make a heavy, full mark, but no stray marks. If you decide to change an answer, erase the unwanted mark very
- There is only one correct answer to each question. Any questions for which more than one response has been blackened w **be counted**.
- Your score is based solely on the number of questions you answer correctly. **It is to your advantage to answer every question.**

1. When 6 M hydrochloric acid is added to an unknown white solid, a colorless gas is produced. What is a possible identity for this solid?

(A) calcium nitrate **(B)** copper(II) chloride

(C) potassium sulfate **(D)** sodium carbonate

2. What is the first change that occurs when $I_2(s)$ is heated slowly at one atmosphere pressure?

(A) The solid melts.

(B) The solid vaporizes.

(C) The solid breaks into atoms.

(D) The solid becomes darker in color.

3. Two pure organic compounds melt at 112 ˚C and 114 ˚C, respectively. If equal quantities of them are mixed, at what temperature will the mixture begin to melt?

(A) below 112 ˚C **(B)** at 112 ˚C

(C) between 112 ˚C and 114 ˚C **(D)** above 114 ˚C

4. What is the major reason for using mercury (rather than water) in barometers?

(A) Mercury is much denser than water.

(B) Mercury has a higher boiling point than water.

(C) Mercury is chemically unreactive compared with water.

(D) Mercury expands with a decrease in air pressure; water does not.

5. How can 0.1 g samples of the two white solids, lead(II) chloride and silver chloride, be distinguished from one another?

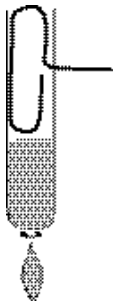
(A) Add 10 mL of cold water to each. The silver chloride will dissolve.

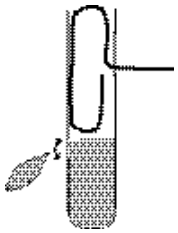
(B) Add 10 mL of hot water to each. The lead(II) chloride will dissolve.

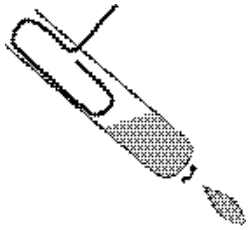
(C) Add 10 mL of sodium chloride to each solution. The lead(II) chloride will become warm and release chlorine gas.

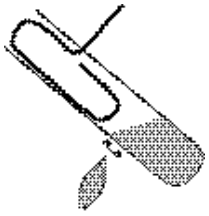
(D) Add 10 mL of zinc chloride solution to each. The silver chloride will change to metallic silver.

6. Which is the proper way to heat a liquid in a test tube?

(A) 

(B) 

(C) 

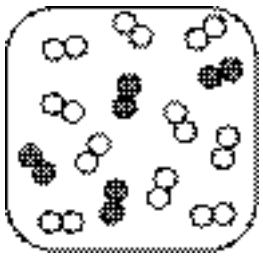
(D) 

7. What is the mass of one molecule of water in grams?

(A) 18 **(B)** 1.1×10^{-21}

(C) 3.0×10^{-23} **(D)** 1.7×10^{-24}

8. If nitrogen atoms are represented as filled circles and and oxygen atoms as open circles, how much NO_2 can be prepared from the mixture shown?



(A) 4 molecules **(B)** 5 molecules

(C) 6 molecules **(D)** 8 molecules

9. A mineral containing iron(II) sulfide but no other sulfides is treated with excess hydrochloric acid to produce hydrogen sulfide. If a 3.15 g sample of the mineral yields 448 mL of hydrogen sulfide gas (measured at 0 ˚C and 760 mm Hg), what is the mass percentage of iron(II) sulfide in the sample?

(A) 20.4 **(B)** 35.5 **(C)** 55.8 **(D)** 71.0

10. Naturally occurring thallium consists of two stable isotopes, Tl-203 and Tl-205 (atomic masses = 203.0 and 205.0, respectively) and has an average atomic mass of 204.4. What is the percentage of Tl-205?

(A) 14.0% **(B)** 30.0% **(C)** 50.0 % **(D)** 70.0 %

11. What is the maximum mass (in grams) of NO that could be obtained from 15.5 g of N_2O_4 and 4.68 g of N_2H_4 when they react? The balanced chemical equation is;

$$2N_2O_4 + N_2H_4 \rightarrow 6NO + 2H_2O.$$

Molar Mass (g·mol^{-1})	
N_2O_4	92.0
N_2H_4	32.0

(A) 4.38 **(B)** 5.04 **(C)** 15.2 **(D)** 26.2

12. What volume of 0.108 M H_2SO_4 is required to neutralize 25.0 mL of 0.145 M KOH?

(A) 16.8 mL **(B)** 33.6 mL

(C) 37.2 mL **(D)** 67.1 mL

13. A gas mixture at 27˚C and 760 mm Hg contains 1.0 g each of He, H_2, N_2 and CO_2. How do their average molecular speeds compare?

(A) $He = H_2 = N_2 = CO_2$ **(B)** $CO_2 < H_2 = N_2 < He$

(C) $He < H_2 < N_2 < CO_2$ **(D)** $CO_2 < N_2 < He < H_2$

14. A 2.00 liter evacuated container has a mass of 1050.0 g. When the container is filled with an unknown gas at 800. mm Hg pressure and 25.0 ˚C the mass is 1052.4 g. What is the molar mass of the gas (in g·mol^{-1})?

(A) 28 **(B)** 31 **(C)** 54 **(D)** 56

15. A mixture of 0.100 mol of N_2 and 0.200 mol of O_2 is collected over H_2O at an atmospheric pressure of 750. mm Hg and a temperature of 22 ˚C. What is the partial pressure (in mmHg) of O_2 in this mixture?

Vapor pressure @ 22 ˚C	
H_2O	22 mmHg

(A) 478 **(B)** 485 **(C)** 500 **(D)** 515

16. A solid is insoluble in water, does not conduct electricity, and does not melt below 1000 ˚C. This solid could be

(A) Pt. **(B)** SiC. **(C)** CsCl. **(D)** $C_{10}H_{22}$.

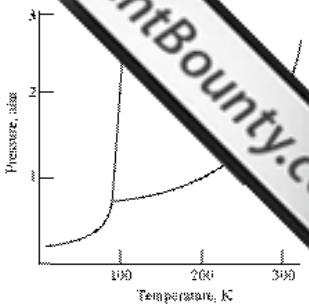
17. Which property(ies) of a liquid increases when the temperature is raised?

I. vapor pressure
II. surface tension

(A) I only **(B)** II only

(C) Both I and II **(D)** Neither I nor II

18. Which statement is correct about the substance represented by this phase diagram?



(A) The solid sublimes at 1 atm pressure.

(B) Its normal boiling point is above 300K.

(C) It exists as a liquid at 25 ˚C and 1 atm pressure.

(D) The density of the solid is greater than that of the liquid.

19. Which equation represents the reaction for the standard enthalpy of formation, ΔH_f°, for $B_5H_9(g)$ at 298 K and 1 atm?

(A) $5B(s) + 9H(g) \rightarrow B_5H_9(g)$

(B) $2B(s) + 3BH_3(g) \rightarrow B_5H_9(g)$

(C) $\frac{5}{2}B_2(g) + \frac{9}{2}H_2(g) \rightarrow B_5H_9(g)$

(D) $5\,B(s) + \frac{9}{2}H_2(g) \rightarrow B_5H_9(g)$

20. $C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$ $\Delta H^\circ = -1427.7$ kJ

If the enthalpy of vaporization for $H_2O(l)$ is 44.0 kJ/mol, what is ΔH° for this reaction if $H_2O(l)$ is formed instead of $H_2O(g)$?

(A) –1295.7 kJ **(B)** –1383.7 kJ

(C) –1471.7 kJ **(D)** –1559.7 kJ

21. A gold ring that weighs 3.81 g is heated to 84.0 ˚C and placed in 50.0 g of H_2O at 22.1 ˚C. What is the final temperature?

Specific Heat Capacity (J·g^{-1}·˚C^{-1})	
Au	0.129
H_2O	4.18

(A) 22.2 ˚C **(B)** 24.0 ˚C

(C) 26.5 ˚C **(D)** 53.1 ˚C

22. Calculate the change in enthalpy, ΔH, for the combustion of 11.2 L of hydrogen gas, measured at 0 ˚C and 1 atm pressure, to form $H_2O(g)$.

ΔH_f° (kJ·mol^{-1})	
$H_2O(g)$	–241.8

(A) –60.5 kJ **(B)** –121 kJ

(C) –484 kJ **(D)** –2710 kJ

23. Which reaction proceeds with the greatest increase in entropy?

(A) $H_2(g) + O_2(g) \rightarrow H_2O_2(l)$

(B) $Br_2(l) + F_2(g) \rightarrow 2BrF(g)$

(C) $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$

(D) $4NH_3(g) + 7O_2(g) \rightarrow 4NO_2(g) + 6H_2O(g)$

24. For the reaction,

$$N_2H_4(l) \rightarrow N_2(g) + 2H_2(g) \quad \Delta H^\circ = -50.6 \text{ kJ}.$$

This reaction is

(A) spontaneous at all temperatures.

(B) non-spontaneous at all temperatures.

(C) spontaneous only at low temperatures.

(D) spontaneous only at high temperatures.

25. All of the following are expected to affect the rate of an irreversible chemical reaction EXCEPT

(A) adding a catalyst.

(B) removing some products.

(C) increasing the temperature.

(D) decreasing the reactant concentration.

26. The oxidation of ammonia produces nitrogen and water according to the equation:

$$4NH_3(g) + 3O_2(g) \rightarrow 2N_2(g) + 6H_2O(g)$$

If the rate of formation of N_2 at a certain temperature is 3.0 mol·L^{-1}·s^{-1}, what is the rate of disappearance of O_2?

(A) 2.0 mol·L^{-1}·s^{-1} **(B)** 3.0 mol·L^{-1}·s^{-1}

(C) 4.5 mol·L^{-1}·s^{-1} **(D)** 9.0 mol·L^{-1}·s^{-1}

27. What are the units of the rate constant for a second order reaction when the rate is expressed in mol·L^{-1}·s^{-1}?

(A) s^{-1} **(B)** L·mol^{-1}

(C) L·mol^{-1}·s^{-1} **(D)** L^2·mol^{-2}·s^{-1}

28. For the reaction:

$(CH_3)_3CBr(aq) + OH^-(aq) \rightarrow (CH_3)_3COH(aq) + Br^-(aq)$

it is found that halving the concentration of $(CH_3)_3CBr$ causes the reaction rate to be halved but halving the concentration of OH^- has no effect on the rate. What is the rate law?

(A) Rate = $k\left[(CH_3)_3CBr\right]^{1/2}\left[OH^-\right]$

(B) Rate = $k\left[(CH_3)_3CBr\right]^{2}\left[OH^-\right]$

(C) Rate = $k\left[(CH_3)_3CBr\right]^{1/2}$

(D) Rate = $k\left[(CH_3)_3CBr\right]$

29. A catalyst speeds up a chemical reac

(A) shifting the equilibrium.

(B) increasing the activation energy.

(C) decreasing the reaction enthalpy.

(D) providing an alternate reaction pathway.

30. The activation energy for a reaction can be determined by measuring the reaction rate at different

(A) temperatures.

(B) catalyst concentrations.

(C) reactant concentrations.

(D) times on the reaction curve.

31. $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H < 0$

Which change(s) will increase the quantity of $SO_3(g)$ at equilibrium?

I.	increasing the temperature
II.	reducing the volume of the container
III.	adding He to increase the pressure

(A) I only **(B)** II only

(C) I and III only **(D)** II and III only

32. What is the equilibrium expression for the reaction;

$2ZnS(s) + 3O_2(g) \rightleftharpoons 2ZnO(s) + 2SO_2(g)$

(A) $K = \dfrac{2[SO_2]}{3[O_2]}$ **(B)** $K = \dfrac{[SO_2]^2}{[O_2]^3}$

(C) $K = \dfrac{2[ZnO][SO_2]}{3[ZnS][O_2]}$ **(D)** $K = \dfrac{[ZnO]^2[SO_2]^2}{[ZnS]^2[O_2]^3}$

33. What is the pH of a 0.0015 M solution of HNO_3?

(A) 1.41 **(B)** 2.82 **(C)** 5.65 **(D)** 11.18

34. In a solution of formic acid ($K_a = 1.7\times10^{-4}$), the $[H^+] = 2.3\times10^{-3}$. What is the concentration of formic acid in mol·L^{-1}?

(A) 7.2×10^{-2} **(B)** 3.1×10^{-2}

(C) 5.3×10^{-6} **(D)** 3.9×10^{-7}

35. What is the $[H^+]$ in a solution in which $[HA] = 4.0\times10^{-2}$ and $[A^-] = 2.0\times10^{-2}$. [$K_a = 3.0\times10^{-6}$]

(A) 1.5×10^{-6} **(B)** 3.0×10^{-6}

(C) 6.0×10^{-6} **(D)** 3.8×10^{-3}

36. When the compounds below are arranged in order of increasing solubility in water, which order is correct?

	K_{sp}
$BaCO_3$	2.6×10^{-9}
$BaSO_4$	1.1×10^{-10}
$CaCO_3$	4.9×10^{-9}
$CaSO_4$	7.1×10^{-9}

(A) $BaCO_3$, $BaSO_4$, $CaCO_3$, $CaSO_4$

(B) $BaSO_4$, $CaCO_3$, $CaSO_4$, $BaCO_3$

(C) $CaSO_4$, $CaCO_3$, $BaCO_3$, $BaSO_4$

(D) $BaSO_4$, $BaCO_3$, $CaCO_3$, $CaSO_4$

37. Which represents an oxidation?

(A) $BrO^- \rightarrow Br_2$ **(B)** $NO_2 \rightarrow N_2O_4$

(C) $Cr^{3+} \rightarrow CrO_4^{2-}$ **(D)** $VO_3^- \rightarrow VO_2^+$

Questions 38 and 39 should be answered using the unbalanced equation;

$$ClO_3^- + Br^- \rightarrow Cl_2 + Br_2$$

38. Which is the reducing agent?

(A) ClO_3^- **(B)** Br^- **(C)** Cl_2 **(D)** Br_2

39. When this equation is balanced, what is the Br^- / ClO_3^- ratio?

(A) 1/1 **(B)** 2/1 **(C)** 3/1 **(D)** 5/1

Questions 40 and 41 require the use of these Standard Reduction Potentials.

Half Reaction	E^o (V)
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	–0.763
$Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$	–0.408
$Tl^+(aq) + e^- \rightarrow Tl(s)$	–0.336
$Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$	+0.161
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.769

40. Use the standard reduction potentials to find the standard cell potential, E°, for the reaction;

$$Zn(s) + 2Tl^+(aq) \rightarrow Zn^{2+}(aq) + 2Tl(s)$$

(A) 0.427 V **(B)** 0.091 V

(C) –0.091 V **(D)** –0.427 V

41. Based on the standard reduction potentials above, which reaction(s) is(are) spontaneous?

I. $Cr^{2+}(aq) + Fe^{3+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{2+}(aq)$

II. $Cu^{2+}(aq) + Fe^{2+}(aq) \rightarrow Cu^+(aq) + Fe^{3+}(aq)$

(A) I only **(B)** II only

(C) Both I and II **(D)** Neither I nor II

42. Solutions of $AgNO_3$, $CuSO_4$ and $AuCl_3$ are electrolyzed in the apparatus depicted. If the electrolysis is stopped before any of the ions are deposited completely, how do the number of moles (n) of Ag, Cu and Au deposited compare?

(A) $n_{Ag} = n_{Cu} = n_{Au}$ **(B)** $n_{Ag} < n_{Cu} < n_{Au}$

(C) $n_{Ag} > n_{Cu} > n_{Au}$ **(D)** $n_{Ag} = n_{Cu} > n_{Au}$

43. Which set of quantum numbers (n, ℓ, m_ℓ, m_s) is not permitted by the rules of quantum mechanics?

(A) 1, 0, 0, ½ **(B)** 2, 1, –1, – ½

(C) 3, 3, 1, – ½ **(D)** 4, 3, 2, ½

44. How many unpaired electrons are in a gaseous iron atom in its ground state?

(A) 6 **(B)** 4 **(C)** 2 **(D)** 0

45. Evidence for the electron arrangement in atoms has been obtained primarily from the study of

(A) isotopes. **(B)** radioactivity.

(C) stoichiometry. **(D)** atomic spectra.

46. Which statement about the radii of atoms and their ions is correct?

(A) Cations are smaller than their atoms, anions are larger.

(B) Cations and anions are both smaller than their atoms.

(C) Cations and anions are both larger than their atoms.

(D) Cations are larger than their atoms, anions are smaller.

47. When the atoms; Li, Be, B, Na, are arranged in order of increasing atomic radius which is the correct order?

(A) Li, Be, B, Na **(B)** Li, Na, B, Be

(C) Na, Li, Be, B **(D)** B, Be, Li, Na

48. Which type of radioactive decay produces a daughter nucleus with a higher atomic number?

(A) α **(B)** β^- **(C)** γ **(D)** β^+

49. In which species is the central atom NOT surrounded by exactly 8 valence electrons?

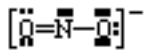
(A) BF_4^- **(B)** NCl_3 **(C)** PCl_4^+ **(D)** SF_4

50. In which pair of formulas are both species polar?

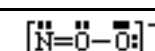
(A) CO_2 and H_2O **(B)** BF_3 and PCl_3

(C) SO_2 and SCl_2 **(D)** CS_2 and NO_2

51. The nitrite ion can be represented;



Which of the structures represent possible resonance forms of this ion?

I. 

II.

(A) I only **(B)** II only

(C) Both I and II **(D)** Neither I nor II

52. Which statement is true about the most stable Lewis structure for CS_2?

(A) There are no lone pairs.

(B) All bonds are double bonds.

(C) The central atom does not have an octet of electrons.

(D) A sulfur atom must be the central atom for the structure to be stable.

53. When the molecules N_2, O_2, F_2 are arranged in order of increasing bond strength, which order is correct?

(A) N_2, O_2, F_2 **(B)** N_2, F_2, O_2

(C) O_2, N_2, F_2 **(D)** F_2, O_2, N_2

54. Which molecule contains three sigma (σ) and two pi (π) bonds?

(A) C_2H_2 **(B)** C_2H_4 **(C)** C_2H_6 **(D)** C_3H_4

55. Which is an isomer of 1-butanol?

(A) 1-propanol **(B)** butanone

(C) 1-chlorobutane **(D)** diethyl ether

56. Which compound is an aldehyde?

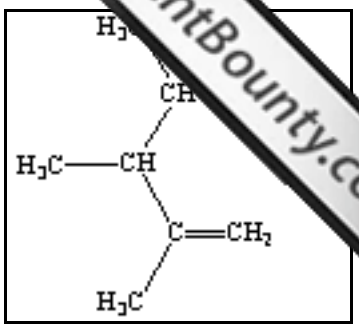
(A) CH_3OCH_3 **(B)** CH_3CHO

(C) CH_3COOH **(D)** CH_3COOCH_3

57. How many isomers exist for dibromobenzene ($C_6H_4Br_2$)?

(A) one **(B)** two **(C)** three **(D)** four

58. What is the name for the compound with the structure?



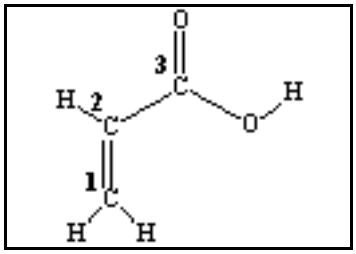
(A) 2-isopropyl-1-butene

(B) 2,3-dimethyl-2-hexene

(C) 2-methyl-3-isopropyl-1-butene

(D) 2,3,4-trimethyl-1-pentene

59. What is the hybridization of carbon atoms 1, 2, and 3, respectively in the structure?



(A) sp^3, sp, sp^2 **(B)** sp^2, sp, sp^2

(C) sp^3, sp^2, sp^2 **(D)** sp^2, sp^2, sp^2

60. When egg white is coagulated the protein is said to be

(A) condensed. **(B)** denatured.

(C) hydrolyzed. **(D)** polymerized.

END OF TEST

Olympiad 2005 Local Section

KEY

Number	Answer
1.	D
2.	B
3.	A
4.	A
5.	B
6.	D
7.	C
8.	D
9.	C
10.	D
11.	C
12.	A
13.	D
14.	A
15.	B
16.	B
17.	A
18.	D
19.	D
20.	D
21.	A
22.	B
23.	B
24.	A
25.	B
26.	C
27.	C
28.	D
29.	D
30.	A

Number	Answer
31.	B
32.	B
33.	B
34.	B
35.	C
36.	D
37.	C
38.	B
39.	D
40.	A
41.	A
42.	C
43.	C
44.	B
45.	D
46.	A
47.	D
48.	B
49.	D
50.	C
51.	A
52.	B
53.	D
54.	A
55.	D
56.	B
57.	C
58.	D
59.	D
60.	B