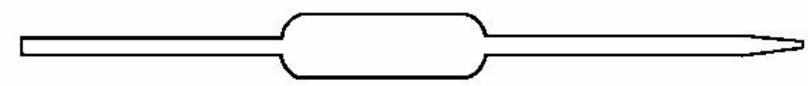
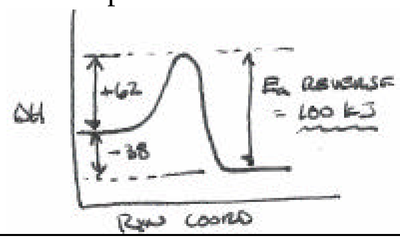
NChO 2006

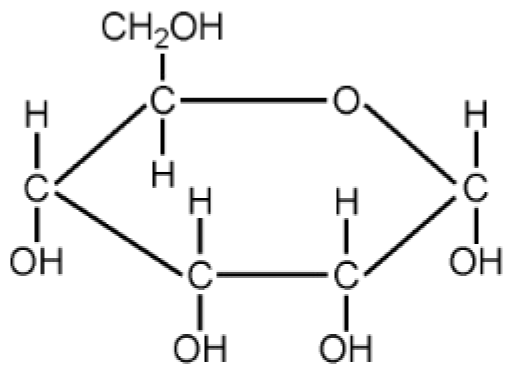
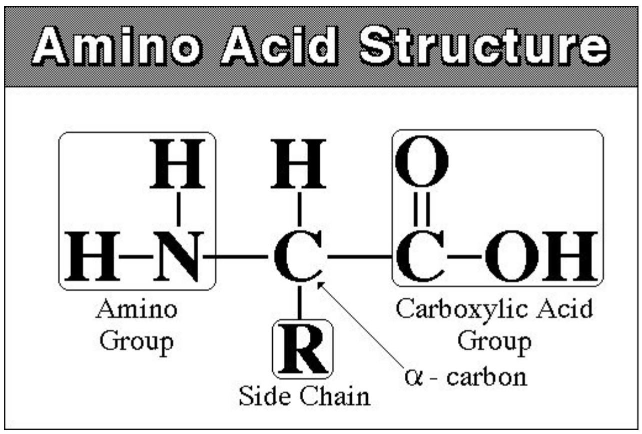
ANNOTATED ANSWERS

1.	**D HCl** H_2 and CH_4 are nonpolar and will dissolve very little in polar water. CO is polar and will dissolve OK, but HCl is very polar and dissolves amazingly well in water. Note: HCl and NH_3 both dissolve extremely well in water.
2.	**C buret** Beakers only have approximate markings. Volumetric pipets and volumetric flasks are only good for measuring exactly 25.0 mL. Here is a picture of a volumetric pipet. 
3.	**B Dehydrating Agent** We did this in class (on the balcony). Sugar's empirical formula is CH_2O. Concentrated H_2SO_4 removes the H_2O leaving C(s). Here is a great movie of the process: http://cwx.prenhall.com/petrucci/medialib/media_portfolio/text_images/060_DehydratSugar.MOV
4.	**C HCl + HNO_3** Gold is a noble metal and does not react with acids, even the oxidizing acids. However, HCl + HNO_3 form Cl_2 in solution which oxidizes the Au $\rightarrow$ Au^{3+} and forms the $AuCl_4^-$ complex ion.
5.	**A 12 g** Read graph carefully. At 80°C solubility = 90 g/100 mL; at 20°C solubility = 60 g/100 mL. Difference = 30 g/100 mL which is 12 g/40 mL
6.	**B too low** An example of a hydrated salt is $CuSO_4 \cdot 5H_2O$. Not heating enough means not all the H_2O is driven off the hydrate. $\therefore$ the mass before heating is fine, but he mass after heating is too large. The mass of H_2O calculated will be too small and the degree of hydration calculated will be too low.
7.	**D $2Ag^+ + S^{2-}$ ® $Ag_2S(s)$** molecular: $2AgNO_3 + Na_2S \rightarrow Ag_2S(s) + 2NaNO_3$ ionic: $2Ag^+ + 2NO_3^- + 2Na^+ + S^{2-} \rightarrow Ag_2S(s) + 2Na^+ + 2NO_3^-$ net ionic: $2Ag^+ + S^{2-} \rightarrow Ag_2S(s)$ Answer is <u>un</u>balanced, but shows the correct species.
8.	**B 1:2** iron(II) oxide = FeO; iron(III) oxide = Fe_2O_3; the mixed oxide, Fe_3O_4 = FeO + Fe_2O_3. This is one Fe^{2+} and two Fe^{3+}
9.	**D 10.2 g** Mass of the products = mass of the reactants. We know C_3H_8, so calculate O_2 reacted and add. $2.20\text{ g } C_3H_8 \times \frac{1\text{ mole } C_3H_8}{44.11\text{ g } C_3H_8} \times \frac{5\text{ mol } O_2}{1\text{ mol } C_3H_8} \times \frac{32.00\text{ g } O_2}{1\text{ mol } O_2}$ = 7.98 g O2; 2.20 g + 7.98 g = **10.18 g**
10.	**D 139 mL** Dilution formula: $V_{before} \cdot \underline{M}_{before} = V_{after} \cdot \underline{M}_{after}$; x (18.0 <u>M</u>) = (25.00 mL) (1.00 <u>M</u>) x = 138.889 mL
11.	**C 5.11%** $6.40\text{ mL AgNO3} \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{0.200\text{ mol } AgNO_3}{1\text{ L}} \times \frac{1\text{ mol } Br^-}{1\text{ mol } AgNO_3} \times \frac{79.9\text{ g } Br^-}{1\text{ mole } Br^-} = 0.102272\text{ g } Br^-$ $\frac{0.102272\text{ g } Br^-}{2.00\text{ g sample}} \times 100 =$ **5.11%**
12.	**B high P/low T** Gases dissolve best at high pressures (Henry's Law) and low temperatures (de-emphasize entropy).

	Entropy favors gaseous state (very random, spread out) rather than dissolved state.
13.	**A HF only** Hydrogen bonding occurs when H is bonded to F, H with N, and H with O.
14.	**C 50.0 g/mol** Molar Mass = grams/moles; 10.00 grams is given. Use PV=nRT to find moles or the fact that this problem is at STP, so 1 mole gas = 22.4 L $\therefore$ 4.48 L = 0.200 moles. Molar mass = $\frac{10.00\text{ g}}{0.200\text{ moles}}$ = **50.0 g/mol**
15.	**A breaking the O-H bond** Evaporating, melting, and subliming are all phase changes. Chemical change (breaking bonds) requires more energy than overcoming intermolecular forces of attraction.
16.	**B** Melting is an orderly lattice becoming less orderly, but still with atoms touching each other. Note: (A) would be evaporation; (C) would be a solid; (D) would be a solid split into pieces.
17.	**D CCl_4** Since the differences in electronegativity between C (2.5) & Cl (3.0) and C (2.5) & H (2.1) are small, these molecules are all non-polar. London forces play the strongest role in determining the boiling temperature. CCl_4 has the largest atoms and the ***most polarizable electron clouds***, $\therefore$ highest BP.
18.	**B** Know how to label the phases in a phase diagram. As you increase pressure, the substance changes from liquid to solid. $\therefore$ solid is more dense than liquid.
19.	**A K_p** The only thing that changes K is temperature.
20.	**A I only** (combustion of charcoal) Which situation gets more random, more disordered, more spread-out. I-yes; II-no (gas $\rightarrow$ liquid); III-no (dissolved ions $\rightarrow$ solid)
21.	**C -40 kJ** D + A $\rightarrow$ 4C desired equation A $\rightarrow$ 2B OK as is ΔH = 40 kJ 2B $\rightarrow$ 2C ΔH x 2 ΔH = -100 kJ D $\rightarrow$ 2C reverse equation $\therefore$ change sign ΔH = +20; 40 + (-100) + 20 = **-40 kJ**
22.	**C 98 kJ** Bond energies of the reactant bonds – bond energies of the product bonds N≡N + 2 H—H $\rightarrow$ N—N + 4 N—H 941 + 2(436) – [159 + 4(389)] = 1813 – 1715 = **98 kJ**
23.	**B 8.63 g** heat lost by hot water = -(heat gained by melting ice + heat gained by ice water warming up) q = mCΔT mass x ΔH_{fus} q = mCΔT; let x = mass of ice (265g)(4.18)(-3.3°C) = -[x·333 + (x)(4.18)(21.70°C) -3655.41 = - (333x + 90.7x) **x = 8.627 g**
24.	**B exothermic reaction with an increase in entropy** Two driving forces, decrease in potential energy (exothermic, $\Delta H<0$) and increase in entropy ($\Delta S>0$)
25.	**D adding 150 mL acid instead of 100 mL acid** Adding more acid does not change the **concentration** of the acid and will not increase the contact between the solution and the solid $CaCO_3$. (A) increases rate by increasing concentration of the acid (B) increases rate by increasing the temperature of the mixture (C) increases rate by increasing surface area (important when there are two different phases involved)
26.	**C 0.18 mol·L^{-1}·min^{-1}**

	Faster rate by a ratio of $\frac{3\text{ mol } N_2}{2\text{ mol } N_2H_4}$
27.	**B Rate = k[A]^2[B]** Expt 1 & 3: [A] constant, [B] doubles, Rate doubles $\therefore$[B]^1 (first order with respect to B) Expt 1 & 2: [B] constant, [A] doubles, Rate x 4 $\therefore$[A]^2 (second order with respect to A)
28.	**D 100 kJ** Draw a picture of the situation. 
29.	**C second** Note: know how to derive the units for these cases and maybe memorize <u>Zero order</u>: Rate = k; k = Rate = mol·L^{-1}·s^{-1} <u>1st order</u>: Rate = k[A]; k = $\frac{\text{Rate}}{[A]}$ = s^{-1} <u>2nd order</u>: Rate = k[A]^2; k = $\frac{\text{Rate}}{[A]^2}$ = L·mol^{-1}·s^{-1}
30.	**A catalyst provides a pathway with a lower activation energy** (B) only temperature changes the equilibrium constant, K (C) increasing temperature increases kinetic energy, not a catalyst (D) the catalyst interacts with the reactants. While it may be true that it interacts with the products, that is not its role.
31.	**A** liquids are not included in the equilibrium expression (their concentrations do not change)
32.	**C increasing the volume of the container** Look for a change that shifts the equilibrium toward the **<u>products</u>**. When the stress is too much room, the equilibrium shifts to the side that takes up more room. The products contain more moles of gas so they take up more room. (A) catalyst speeds up both the forward and reverse reaction so you get to the same equilibrium. (B) endothermic means heat is a reactant. Decreasing the temperature shifts reaction toward the left. (D) slows down both the forward and reverse reaction so you get the same equilibrium.
33.	**A smallest K_a** Note: Be careful of negative exponents. $10^{-5} < 10^{-4}$ and $1.8 < 5.5$
34.	**D pH = 6.30** Note: this question can be a calculation problem or a conceptual problem. This is an ICE box problem with initial values for both HA and A- (common ion effect) You might also notice that this is a buffer. If [HA] = [A$^-$] then pH = pK_a = 6 Since there is some excess conjugate **base** (A$^-$), the buffer's pH will be a little more basic, **6.30**.
35.	**B NH_4Br** pH less than 7 means acidic. Look for an acid or the conjugate acid of a weak base. (A) neutral **(B) conjugate acid of NH_3** (C) conjugate <u>base</u> of HF (D) conjugate <u>base</u> of $HC_2H_3O_2$
36.	**D 6.4 x 10^{-7}M** The problem states that the K_{sp} of MgF_2(s) is 6.4 x 10^{-9}. This implies the equation: MgF_2(s) $\leftrightarrows$ Mg^{2+}(aq) + 2 F^-(aq) The K_{sp} expression is K_{sp} = [Mg^{2+}][F^-]^2 Essentially all of the F^- ions come from the 0.10 $_M$ solution of NaF. Solve for the [Mg^{2+}]. [Mg^{2+}][F^-]^2 = 6.4 x 10^{-9} (x) (0.10)^2 = 6.4 x 10^{-9} x = 6.4 x 10^{-7}
37.	**B VO^{2+} ® VO_3^-** (A) N(+3) $\rightarrow$ N(0) reduction **(B) V(+4) ® V(+5) oxidation**

	(C) Cl(+1) → Cl(-1) reduction (D) Cr(+6) → Cr(+6) no redox change
38.	**D +6** MoO_2Cl_2 Think of this as an algebra problem: x + 2(-2) + 2(-1) = 0 ∴ x = +6
39.	**B $4H^+$** Notice that this is a half-reaction: $2H_2O + S^{2-} \rightarrow SO_2 + 4H^+ + 6e^-$
40.	**B +0.426 V** reduction $E°_{reduction}$ - oxidation $E°_{reduction}$ = -762 – (-.336) = +0.426 V Note: There is no need to <u>double</u> the Tl E° even though you double the equation when you combine the oxidation and reduction reactions. This is because each half cell was compared to the standard H_2 half cell. The differences in electrons involved has already been taken into account.
41.	**A E° +; ΔG° - ; spontaneous** A spontaneous electrochemical cell has a + E° value. A spontaneous reaction has a – ΔG° value.
42.	**C O_2** Anode means oxidation. For electrolysis, either the H_2SO_4 can change or the H_2O will change. Eliminate (A) because $2H^+ + 2e^- \rightarrow H_2$ this is reduction… it would occur only at the <u>cathode</u>. Either H_2O is oxidized: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ (memorize this equation) or SO_4^{2-} is oxidized: S in SO_4^{2-} is already as oxidized as it can be. The water will change.
43.	**D Cs** Cs is a member of the alkali metal family, the most reactive metallic family. Cs is the most reactive of the alkali metals. (Fr would be the most reactive if there were any of it around.)
44.	**C N** N has the most protons in its row and is the smallest atom ∴ the strongest ionization energy.
45.	**A 3f** The first f-orbitals are 4f.
46.	**B 9** Draw the orbital diagram: $1s^2$ $2s^2$ $2p_x^2$ $2p_y^2$ $2p_z^2$ $3s^2$ $3p_x^2$ $3p_y^1$ $3p_z^1$
47.	**B II only** *n* determine energy; ***l* determines the shape (type of orbital)**; m_l determines the orientation
48.	**C lanthanides** lanthanides: the 4f orbitals are being filled actinides: the 5f orbitals are being filled transition metals: the d orbitals are being filled
49.	**A NaF** <u>Small</u> lattice energy, look for <u>large</u> size and <u>small</u> charge. Na^+ and F^-
50.	**A I only (NO)** Draw the Lewis dot structures. NO has 11 electrons, the N atom has 7 electrons. SF_2, all single bonds, each atom has an octet. PF_4^+, four single bonds, each atom has an octet.
51.	**D CH_3CN** Draw the Lewis dot structures. Shortest CN bond is the triple bond in CH_3CN.
52.	**D NO_2^-** Draw the Lewis dot structures. NO_2^- has two resonance structures and a lone pair on the N atom.
53.	**C $C_1 – sp^3$, $C_2 – sp^2$** SN = steric number (# of bonded atoms + # lone pairs); C_1 SN=4 sp^3; C_2 SN=3 sp^2
54.	**D II. & III only ($COCl_2$ & CH_2Cl_2)** Draw the Lewis dot structures. CO_2 is linear and non-polar. $COCl_2$ and CH_2Cl_2 are not symmetrical and polar.
55.	**B benzene** This is just a definition: benzene compounds are called aromatic compounds.
56.	**D 1,2-dichloroethene** Geometric isomers mean cis-/trans- isomers. Look for a double bond (-ene) with groups on the two carbons of the double bond.

57.	**B CH_3COOH (carboxylic acid)** You need to recognize that an acidified solution of $Cr_2O_7^{2-}$ is a strong oxidizer. An alcohol can be oxidized into an aldehyde and then further oxidized into a carboxylic acid.
58.	**A -O-H** Simple sugars have the empirical formula CH_2O. Glucose is shown. The -O-H bonds allow hydrogen bonding and solubility in H_2O.  Source: http://www2.glos.ac.uk/GDN/origins/images/sugar.gif
59.	**C esters** Esters are used for flavorings and fragrances. A standard high school lab is making Oil of Wintergreen from methyl alcohol and salicylic acid.
60.	**B nitrogen**  Source: http://www.hcc.mnscu.edu/programs/dept/chem/V.27/amino_acid_structure_2.jpg