Chemistry 11 Exam 1Spring 2006

When answering questions be sure to write clearly. Be sure to show your thinking as you answer the questions. You must show work in order to receive credit for your answers. For any answers greater than 1000 or less than 0.01 be sure to report the answer using scientific notation or metric prefixes. All answers should include units where appropriate and the correct number of significant figures.

1. Name the following compounds: 6 points

$FeCr_2O_7$ ________________________________

SiS_2 ________________________________

K_2O ________________________________

Write the formulas for the following compounds:

triphosphorous hexafluoride ____________________

cobalt(II) carbonate ____________________

calcium fluoride ____________________

2. Naturally occurring carbon exists in two isotopic forms, carbon-12 and carbon-13. The atomic masses and natural abundances of each isotope are: 12.0000u (98.93%) and 13.00335u (1.07%). Use this data to calculate the atomic mass of carbon. Then write the isotopic notation for each isotope. 6 points

3. How many moles of carbon are in 2.55×10^{21} atoms of carbon? 2 points

How many moles of carbon are there in 6.25×10^{21} molecules of glucose, $C_6H_{12}O_6$? 2 points

How many carbon atoms are in 115.00 g of glucose, $C_6H_{12}O_6$? 2 points

4. The systematic name for AZT, a drug used to treat patients with HIV/AIDS, is 1-[4-azido-5-(hydroxymethyl)oxolan-2-yl]-5-methyl-pyrimidine-2,4-dione. The mass percentage composition of AZT is:

44.94% C 4.90% H 26.21% N 23.95% O

What is the empirical formula of AZT? 6 points

5. Balance the following chemical equations using the smallest whole numbers possible: 4 points

$C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$ ______________________________

$Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$ ______________________________

6. Consider the following equation: $\underset{\text{octane}}{2C_8H_{18}(l)} + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

If 12 molecules of octane are mixed with an excess of oxygen gas how many molecules of water will be produced? 2 points.

If 8 molecules of octane are mixed with 75 molecules of oxygen gas how many molecules of carbon dioxide will be produced? 2 points

If 12.50 g of octane (molar mass = 114.23 g/mol) is mixed with excess oxygen, how many moles of carbon dioxide will be produced? 2 points

If 14.050 g of water are produced in the reaction, how many grams of octane were used? 2 points

7. Complete the following chemical equations. Balance the equations. Then write the complete ionic equations and the net ionic equations: 12 points

$Co(NO_3)_3(aq) + NaOH(aq) \rightarrow$

Complete ionic:

Net ionic:

$HNO_3(aq) + KOH(aq) \rightarrow$

Complete ionic:

Net ionic:

8. A solution of nitric acid (HNO_3) in water is prepared. In this solution 2.500 g of nitric acid is dissolved in 50.00 mL of water. What is the molar concentration of the solution? 4 points

A 45.00 mL sample of 1.25 M HNO_3 is diluted to 200.00 mL. What is the molarity of the resulting solution? 4 points

9. Rank the following solutions in the order of highest electrical conductivity to lowest. Then briefly explain your ranking. 4 points

Use: **1** for highest, **2** for next highest, **3** for next highest, and **4** for lowest.

a. 0.25 moles NaCl added to 2.00 L of water____

b. 0.50 moles NaCl added to 2.00 L of water____

c. 0.50 moles NaCl added to 1.00 L of water____

d. 1.0 moles AgCl added to 1.00 L of water____

$N_A = 6.022 \times 10^{23}$/mol

Solubility rules: Alkali metals, ammonium, nitrate, perchlorate (ClO_4^-), chlorate (ClO_3^-), nitrate, and acetate ($C_2H_3O_2^-$) salts are soluble.

Chlorides, bromides (Br^-), and iodides (I^-) are soluble unless combined with Ag^+, Pb^{+2}, or Hg_2^{+2}.

Sulfates are soluble unless combined with Pb^{+2}, Ca^{+2}, Sr^{+2}, Ba^{+2}, or Hg_2^{+2}

Metal hydroxides are insoluble unless combined with alkali metals, Ca^{+2}, Sr^{+2}, or Ba^{+2}.

Phosphates, carbonates, sulfates, and sulfites are insoluble unless combined with alkali metals or ammonium cations.

Metric prefixes
Nano (n) = 10^{-9}
Micro (μ) = 10^{-6}
Milli (m) = 10^{-3}
Centi (c) = 10^{-2}
Kilo (k) = 10^{3}
Mega (M) = 10^{6}
Giga (G) = 10^{9}

Chemistry 11 Spring 2006 Exam 1

When answering questions be sure to write clearly. Be sure to show your thinking as you answer the questions. You must show work in order to receive credit for your answers. For any answers greater than 1000 or less than 0.01 be sure to report the answer using scientific notation or metric prefixes. All answers should include units where appropriate and the correct number of significant figures.

1. Name the following compounds:

$FeCr_2O_7$ iron(II) dichromate

SiS_2 silicon disulfide

K_2O potassium oxide

Write the formulas for the following compounds:

triphosphorous hexafluoride P_3F_6

cobalt(II) carbonate $CoCO_3$

calcium fluoride CaF_2

2. Naturally occurring carbon exists in two isotopic forms, carbon-12 and carbon-13. The atomic masses and natural abundances of each isotope are: 12.0000u (98.93%) and 13.00335u (1.07%). Use this data to calculate the atomic mass of carbon. Then write the isotopic notation for each isotope.

average mass = 0.9893(12.0000u) + 0.0107(13.00335u) = 12.01 u

$^{12}_{6}C$ $^{13}_{6}C$

3. How many moles of carbon are in 2.55×10^{21} atoms of carbon?

$(2.55 \times 10^{21})(1 \text{ mol}/6.022 \times 10^{23}) = 4.23 \times 10^{-3}$ mol

How many moles of carbon are there in 6.25×10^{21} molecules of glucose, $C_6H_{12}O_6$?

$(6.25 \times 10^{21}$ molecules$)(1 \text{ mol}/6.022 \times 10^{23})(6$ carbon atoms/1 molecule$) = 0.0623$ mol

How many carbon atoms are in 115.00 g of glucose, $C_6H_{12}O_6$?

115.00 g (1 mol $C_6H_{12}O_6$/180.16 g)(6.022×10^{23} $C_6H_{12}O_6$/1 mol)(6 C atoms/1 molecule) = 2.306×10^{24} carbon atoms

4. The systematic name for AZT, a drug used to treat patients with HIV/AIDS, is 1-[4-azido-5-(hydroxymethyl)oxolan-2-yl]-5-methyl-pyrimidine-2,4-dione. The mass percentage composition of AZT is:

44.94% C 4.90% H 26.21% N 23.95% O

What is the empirical formula of AZT?

44.94 g C(1 mol C/12.011 g C) = 3.742 mol C

4.90 g H (1 mol H/1.008 g H) = 4.861 mol H

26.21 g N(1 mol N/14.01 g N) = 1.871 mol N

23.95 g O(1 mol O/15.999 g O) = 1.497 mol O

C: 3.742/1.497 = 2.5

H: 4.861/1.497 = 3.25

N: 1.871/1.497 = 1.25

O: 1.497/1.497 = 1

Multiply by 4 to give whole numbers:

C: 2.5 x 4 = 10

H: 3.25 x 4 = 13

N: 1.25 x 4 = 5

O: 1 x 4 = 4

$C_{10}H_{13}N_5O_4$

5. Balance the following chemical equations using the smallest whole numbers possible:

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

6. Consider the following equation: $\underset{\text{octane}}{2C_8H_{18}(l)} + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

If 12 molecules of octane are mixed with an excess of oxygen gas how many molecules of water will be produced?

12 molecules C_8H_{18}(18 molecules H_2O/2 molecules C_8H_{18}) = 108 molecules H_2O

If 4 molecules of octane are mixed with 40 molecules of oxygen gas how many molecules of carbon dioxide will be produced?

8 molecules C_8H_{18}(16 molecules CO_2/2 molecules C_8H_{18}) = 64 molecules CO_2

75 molecules O_2(16 molecules CO_2/25 molecules O_2) = **48 molecules CO_2**

If 12.50 g of octane (molar mass = 114.23 g/mol) is mixed with excess oxygen, how many moles of carbon dioxide will be produced?

12.50 g C_8H_{18}(1 mol C_8H_{18}/114.23 g C_8H_{18})(16 mol CO_2/2 mol C_8H_{18}) = 0.875 mol CO_2

If 14.050 g of water is produced in the reaction, how many grams of octane were used?

14.050 g H_2O(1 mol H_2O/18.016 g H_2O)(2 mol C_8H_{18}/18 mol H_2O)(114.23 g C_8H_{18}/1 mol C_8H_{18}) = 9.898 g C_8H_{18}

7. Complete the following chemical equations. Include states of matter. Balance the equations. Then write the complete ionic equations and the net ionic equations:

$Co(NO_3)_3(aq) + 3NaOH(aq) \rightarrow Co(OH)_3(s) + 3NaNO_3(aq)$

Complete ionic:

$Co^{+3}(aq) + 3NO_3\text{-}(aq) + 3Na^+(aq) + 3OH\text{-}(aq) \rightarrow Co(OH)_3(s) + 3Na_+(aq) + 3NO_3\text{-}(aq)$

Net ionic:

$Co^{+3}(aq) + 3OH^-(aq) \rightarrow Co(OH)_3(s)$

$HNO_3(aq) + KOH(aq) \rightarrow H_2O(l) + KNO_3(aq)$

Complete ionic: $H^+(aq) + NO_3\text{-}(aq) + K^+(aq) + OH\text{-}(aq) \rightarrow H_2O(l) + K^+(aq) + NO_3\text{-}(aq)$

Net ionic: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

8. A solution of nitric acid (HNO_3) in water is prepared. In this solution 2.500 g of nitric acid is dissolved in 50.00 mL of water. What is the molar concentration of the solution?

2.500 g HNO_3(1 mol HNO3/63.018 g) = 0.03967 mol HNO_3

$M = n/V = (0.03967 \text{ mol}/50 \times 10^{-3} \text{ L}) = 0.7934 \text{ mol/L}$

A 45.00 mL sample of 1.25 M HNO_3 is diluted to 200.00 mL. What is the molarity of the resulting solution?

$n = MV = 1.25\text{mol/L}(45.00 \times 10^{-3}\text{L}) = 0.05625$ mol HNO_3

$M = n/V = 0.05625 \text{ mol}/200 \times 10^{-3}\text{L} = 0.281 \text{ mol/L}$

9. Rank the following solutions in the order of highest electrical conductivity to lowest. Then explain your ranking.

a. 0.25 moles NaCl added to 2.00 L of water.(3) Third highest concentration of a strong electrolyte.
b. 0.50 moles NaCl added to 2.00 L of water.(2) Second highest concentration of a strong electrolyte.
c. 0.50 moles NaCl added to 1.00 L of water.(1) Highest concentration of a strong electrolyte.
d. 1.0 moles AgCl added to 1.00 L of water.(4) Not soluble in water. Very few ions will be dissolved to conduct.

General Chemistry 1
Exam 2
Spring 2006 May 11, 2006 Section D01B

There are 20 questions in this exam. Answer all 20, showing your reasoning where possible. Each question is valued at 5 points. Be sure to include units when reporting numerical answers. Pay attention to significant figures as well. The final page contains equations and constants that you can use.

1. Assign oxidation numbers to each of the elements in the following reaction:
$Ag(s) + NO_3^-(aq) \rightarrow NO_2(g) + Ag^+(aq)$

— — — — — —

Which reactant was oxidized? Explain.____

____

Which reactant was reduced? Explain.____

____

2. Balance the following reaction in acidic solution.
$HNO_2(aq) + MnO_4^-(aq) \rightarrow Mn^{+2}(aq) + NO_3^-(aq)$

3. What is the physical meaning of a negative value for heat?_________________________

____

4. How much heat in kilojoules is required to bring 250. g of water 15°C to 99°C?

5. Nitrogen dioxide reacting with water produces ammonia and oxygen:
$4NO_2(g) + 6H_2O(g) \rightarrow 4NH_3(g) + 7O_2(g)$ $\Delta H^o = +1132$ kJ

a. Calculate the amount of energy for the reaction of 1 mol of $NO_2(g)$?

b. Calculate the amount of energy for the reaction of 15.55 g of water.

6. Which of the following equations has a value of ΔH that would be properly labeled as $\Delta H°_f$? Explain.

a. $2Fe(s) + O_2(g) \rightarrow 2FeO(s)$

b. $SO_2(g) + \frac{1}{2}O_2(g) \rightarrow SO_3(g)$

c. $N_2(g) + \frac{5}{2}O_2(g) \rightarrow N_2O_5(g)$

____

____

7. The thermochemical equation for the reaction of ethylene gas is shown:
$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
Use the thermodynamic data provided at the end of this exam to calculate the heat of reaction for this reaction.

8. Use equations 1 and 2 to determine the heat of reaction for:

$$CuO(s) + CO(g) \rightarrow Cu(s) + CO_2(g)$$

1. $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ $\qquad \Delta H° = -566.1 \text{ kJ}$
2. $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$ $\qquad \Delta H° = -310.5 \text{ kJ}$

9. What is the Pauli exclusion principle? What effect does it have on the populating of orbitals by electrons?

____

____

____

10. A photon has a wavelength of 340 nm.
a. What is the frequency of the photon?

b. What is the energy of the photon?

11. The outermost electron in an atom has the following quantum numbers:
$n=3$, $l=2$, $m_l=-2$, $m_s=-1/2$

a. What is the name of the orbital that the electron occupies?

b. What kind of element is the atom? Explain.

Alkali metal **alkaline earth metal** **transition metal** **nonmetal**

____

____

12. Write the electron configuration (spdf) for iron (Fe). Then show the shorthand notation based on noble gas configurations.

13. The atomic radii for the first four alkali metals are shown below. Explain this pattern.

____

____

____

Metal	radius (pm)
Li	152
Na	186
K	227
Rb	248

14. Write the Lewis structure for the following species:

CF_4 **CO_2**

15. Show, by drawing Lewis structures, how the nitrite ion (**NO_2^-**) possesses two resonance forms.

16. Use VSEPR Theory to determine the geometric structure (name and bond angles) for:

CH_4 **$SbCl_6^-$**

17. Explain the general idea behind hybrid orbitals.

____

____

18. Consider the following molecule. Label the hybridization for each carbon atom in the structure.

$$\overset{1}{H_2C}=\overset{2}{CH}-\overset{3}{CH_3}$$

carbon 1___________

carbon 2___________

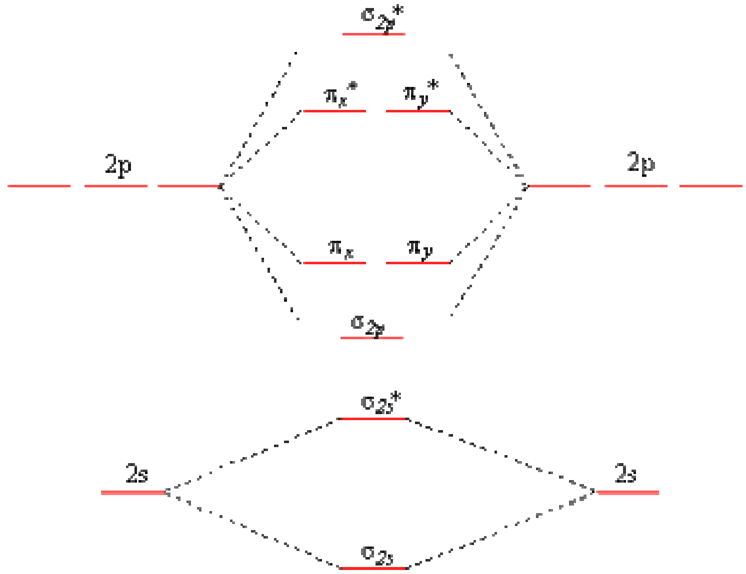
carbon 3___________

19. Explain the general idea behind molecular orbitals.

____

____

20. Fill in the electronic configuration for molecular oxygen (O_2) using the molecular orbital energy diagram



given below.

b. What is the bond order for the molecule? ____________

c. Is the molecule paramagnetic or diamagnetic? Explain.__________________________

____

$\Delta E = q + w$ $\Delta H = q$ $q = -ms\Delta T$ $C = ms$

$\Delta H°_{rxn} = \sum n_p \Delta H°_p - \sum n_r \Delta H°_r$ $T_K = T_C + 273$

$h = 6.626 \times 10^{-34}$ Js $c = 3.00 \times 10^8$ m/s $E = h\nu = hc/\lambda$

atto (a) = 10^{-18}	femto (f) = 10^{-15}	pico (p) = 10^{-12}	nano (n)= 10^{-9}
micro (μ)= 10^{-6}	milli (m) = 10^{-3}	centi (c)= 10^{-2}	kilo (k) = 10^{3}
mega (M) = 10^{6}	giga (G) = 10^{9}	tera (T) = 10^{12}	peta (P) = 10^{15}

specific heat capacities

substance	**s** $\frac{J}{g°C}$
graphite	0.711
copper	0.387
ethyl alcohol	2.45
gold	0.129
granite	0.803
iron	0.445
lead	0.128
water (l)	4.18

Heats of formation (kJ/mol)

CO(g)	-110	CO_2(g)	-394
CH_4(g)	-74.9	C_2H_2 (g)	+227
C_2H_4(g)	+51.9	C_2H_6(g)	-84.5
C_3H_8(g)	-104	C_4H_{10}(g)	-126
H_2O(g)	-241.8	H_2O(l)	-285.9
H_2O_2(l)	-187.8		

1 **H** Hydrogen 1.00794																	2 **He** Helium 4.003
3 **Li** Lithium 6.941	4 **Be** Beryllium 9.012182											5 **B** Boron 10.811	6 **C** Carbon 12.0107	7 **N** Nitrogen 14.00674	8 **O** Oxygen 15.9994	9 **F** Fluorine 18.9984032	10 **Ne** Neon 20.1797
11 **Na** Sodium 22.989770	12 **Mg** Magnesium 24.3050											13 **Al** Aluminum 26.981538	14 **Si** Silicon 28.0855	15 **P** Phosphorus 30.973761	16 **S** Sulfur 32.066	17 **Cl** Chlorine 35.4527	18 **Ar** Argon 39.948
19 **K** Potassium 39.0983	20 **Ca** Calcium 40.078	21 **Sc** Scandium 44.955910	22 **Ti** Titanium 47.867	23 **V** Vanadium 50.9415	24 **Cr** Chromium 51.9961	25 **Mn** Manganese 54.938049	26 **Fe** Iron 55.845	27 **Co** Cobalt 58.933200	28 **Ni** Nickel 58.6934	29 **Cu** Copper 63.546	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.723	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.92160	34 **Se** Selenium 78.96	35 **Br** Bromine 79.904	36 **Kr** Krypton 83.80
37 **Rb** Rubidium 85.4678	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.90585	40 **Zr** Zirconium 91.224	41 **Nb** Niobium 92.90638	42 **Mo** Molybdenum 95.94	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.90550	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.8682	48 **Cd** Cadmium 112.411	49 **In** Indium 114.818	50 **Sn** Tin 118.710	51 **Sb** Antimony 121.760	52 **Te** Tellurium 127.60	53 **I** Iodine 126.90447	54 **Xe** Xenon 131.29
55 **Cs** Cesium 132.90545	56 **Ba** Barium 137.327	57 **La** Lanthanum 138.9055	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.9479	74 **W** Tungsten 183.84	75 **Re** Rhenium 186.207	76 **Os** Osmium 190.23	77 **Ir** Iridium 192.217	78 **Pt** Platinum 195.078	79 **Au** Gold 196.96655	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.3833	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.98038	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
87 **Fr** Francium (223)	88 **Ra** Radium (226)	89 **Ac** Actinium (227)	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (263)	107 **Bh** Bohrium (262)	108 **Hs** Hassium (265)	109 **Mt** Meitnerium (266)	110 (269)	111 (272)	112 (277)	113	114				

58 **Ce** Cerium 140.116	59 **Pr** Praseodymium 140.90765	60 **Nd** Neodymium 144.24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.964	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.92534	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.93032	68 **Er** Erbium 167.26	69 **Tm** Thulium 168.93421	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.967
90 **Th** Thorium 232.0381	91 **Pa** Protactinium 231.03588	92 **U** Uranium 238.0289	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

General Chemistry 1
Exam 2
Spring 2006 May 11, 2006 Section D01B

There are 20 questions in this exam. Answer all 20, showing your reasoning where possible. Each question is valued at 5 points. Be sure to include units when reporting numerical answers. Pay attention to significant figures as well. The final page contains equations and constants that you can use.

1. Assign oxidation numbers to each of the elements in the following reaction:
$Ag(s) + NO_3^-(aq) \rightarrow NO_2(g) + Ag^+(aq)$
0 +5 -2 +4 -2 +1

Which reactant was oxidized? Explain. Silver started with an oxidation number of 0 and ended with an oxidation number of +1. This represents a loss of an electron so the silver was oxidized.

Which reactant was reduced? Explain. Nitrogen in the nitrate ion has an oxidation number of +5. The oxidation of nitrogen in nitrogen dioxide is +4. So nitrogen was reduced (an electron was gained). THe reactant is nitrate so we can also say that nitrate was reduced.

2. Balance the following reaction in acidic solution.
$HNO_2(aq) + MnO_4^-(aq) \rightarrow Mn^{+2}(aq) + NO_3^-(aq)$

$HNO_2 \rightarrow NO_3^-$	$MnO_4^- \rightarrow Mn^{+2}$
$H_2O + HNO_2 \rightarrow NO_3^-$	$MnO_4- \rightarrow 4H_2O + Mn^{+2}$
$H_2O + HNO_2 \rightarrow NO_3^- + 3H^+$	$8H^+ + MnO_4^- \rightarrow 4H_2O + Mn^{+2}$
$H_2O + HNO_2 \rightarrow NO_3^- + 3H^+ + 2e^-$	$5e^- + 8H^+ + MnO_4^- \rightarrow 4H_2O + Mn^{+2}$
$5(H_2O + HNO_2 \rightarrow NO_3^- + 3H^+ + 2e^-)$	$2(5e^- + 8H^+ + MnO_4^- \rightarrow 4H_2O + Mn^{+2})$
$5H_2O + 5HNO_2 \rightarrow 5NO_3^- + 15H^+ + 10e^-$	$10e^- + 16H^+ + 2MnO_4^- \rightarrow 8H_2O + 2Mn^{+2}$

$\mathbf{5HNO_2 + H^+ + 2MnO_4^- \rightarrow 5NO_3^- + 2Mn^{+2} + 3H_2O}$

3. What is the physical meaning of a negative value for heat? Negative heat means that heat was released in a process. This is called an exothermic process.

4. How much heat (in kilojoules) is required to bring 250. g of water 15°C to 99°C?

$q = ms\Delta T$

$q = (250.\text{g})(4.18\frac{\text{J}}{\text{g°C}})(99°\text{C}-15°\text{C}) = 8.78 \times 10^4\ \text{J} = 87.8\ \text{kJ}$

5. Nitrogen dioxide reacting with water produces ammonia and oxygen:
$4NO_2(g) + 6H_2O(g) \rightarrow 4NH_3(g) + 7O_2(g)$ $\Delta H^o = +1132\ \text{kJ}$

a. Calculate the amount of energy for the reaction of 1 mol of $NO_2(g)$?

$\Delta H^o = \frac{+1132\ \text{kJ}}{4\,\text{mol}\,NO_2}(1\,\text{mol}\,NO_2) = 283\,\text{kJ}$

b. Calculate the amount of energy for the reaction of 15.55 g of water.

$(15.55\,\text{g}\,H_2O)(\frac{1\,\text{mol}\,H_2O}{18.016\,\text{g}\,H_2O}) = 0.863\,g\,H_2O$

$\Delta H^o = \frac{+1132\ \text{kJ}}{6\,\text{mol}\,H_2O}(0.863\,\text{mol}\,H_2O) = 163\,\text{kJ}$

6. Which of the following equations has a value of ΔH that would be properly labeled as $\Delta H°_f$? Explain.

a. $2Fe(s) + O_2(g) \rightarrow 2FeO(s)$

b. $SO_2(g) + \frac{1}{2} O_2(g) \rightarrow SO_3(g)$

c. $N_2(g) + \frac{5}{2} O_2(g) \rightarrow N_2O_5(g)$

Strictly speaking, **c** only because heat of formation is the amount of energy associated with forming 1 mole of a compound from the constituent elements. a is the formation of 2 moles of a compound. b is the formation of a compound from an element and a compound.

7. The thermochemical equation for the reaction of ethylene gas is shown:
$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
Use the thermodynamic data provided at the end of this exam to calculate the heat of reaction for this reaction.

$\Delta H°_{rxn} = \sum n_p \Delta H°_p - \sum n_r \Delta H°_r =$

2 mol CO_2(-394 kJ/mol) + 2 mol H_2O(-241.8 kJ/mol) - 1 mol C_2H_4(+51.9 kJ/mol) - 0
= **-1320. kJ**

8. Use equations 1 and 2 to determine the heat of reaction for:
$CuO(s) + CO(g) \rightarrow Cu(s) + CO_2(g)$

1. $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ $\Delta H° = -566.1$ kJ
2. $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$ $\Delta H° = -310.5$ kJ

Divide equation 1 by 2: -566.1 kJ/2 = = -283.1 kJ
Reverse equation equation 2 and divide by 2: -(-310.5 kJ)/2 = +155.3 kJ
Overall: -283.1 kJ + 155.3 kJ = -127.8 kJ

9. What is the Pauli exclusion principle? What effect does it have on the populating of orbitals by electrons?

No two electrons can simulataneously possess identical values for the 4 quantum numbers in a single atom. The result of this fact is that electrons in multielectron atoms populate orbitals at different energies.

10. A photon has a wavelength of 340 nm.
a. What is the frequency of the photon? b. What is the energy of the photon?

$$\upsilon\lambda = c = 3.0x10^8 \frac{m}{s}$$

$$\upsilon = \frac{c}{\lambda} = \frac{3.0x10^8 \frac{m}{s}}{340x10^{-9} m} = 8.8x10^{14} s^{-1} (Hz)$$

$$E = h\upsilon = \frac{hc}{\lambda} = 6.626x10^{-34} J \cdot s (8.8x10^{14} s^{-1}) = 5.8x10^{-19} J$$

11. The outermost electron in an atom has the following quantum numbers:
n=3, l=2, m_l=-2, m_s=-1/2

a. What is the name of the orbital that the electron occupies? **3d**
b. What kind of element is the atom? Explain.

Alkali metal **alkaline earth metal** **transition metal** **nonmetal**

Transition metal. When an outermost electron has an l value of 2 then the electron is in a d orbital. The d orbitals are the outermost occupied orbitals in transition metals.

12. Write the electron configuration (spdf) for iron (Fe). Then show the shorthand notation based on noble gas configurations.
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$ = [Ar] $4s^2\ 3d^6$

13. The atomic radii for the first four alkali metals are shown below. Explain this pattern.

Moving down the periodic table the outermost occupied orbitals increase (n= 1,2,3,4,5,...). As the outermost occupied orbital increases for n the size of the orbital also increases. For Na the outermost occupied orbital is n=3. This orbital is larger than for n=2 as found for the outermost occupied orbital in Li.

14. Write the Lewis structure for the following species:
CF_4 C is central atom, bonded to 4 surrounding F atoms. Each F atom has 3 lone pairs.
CO_2 O=C=O with 2 lone pairs on each oxygen atom

15. Show, by drawing Lewis structures, how the nitrite ion (**NO_2^-**) possesses two resonance forms. (18 electrons total).

	O_A-N=O_B]$^{-1}$	<--->	O_A=N-O_B]$^{-1}$
O_A:	3 lone pairs		2 lone pairs
N:	1 lone pair		1 lone pair
O_B:	2 lone pairs		3 lone pairs

16. Use VSEPR Theory to determine the geometric structure (name and bond angles) for:

CH_4
tetraheadral (109.5°)

$SbCl_6^-$
octahedral (90°)

17. Explain the general idea behind hybrid orbitals.
Atomic orbitals with the same principal quantum number combine to form new orbitals that possess energies that are intermediate compared to the energies of the atomic orbitals. These hybridized orbitals overlap with orbitals from other atoms to form covalent bonds.

18. Consider the following molecule. Label the hybridization for each carbon atom in the structure.

$\overset{1}{H_2C}=\overset{2}{C}H-\overset{3}{C}H_3$

carbon 1 sp^2 carbon 2 sp^2 carbon 3 sp^3

19. Explain the general idea behind molecular orbitals.
When atoms form covalent bonds each can provide one electron. In molecular orbitals each atom possesses an atomic orbital. These two atomic orbitals combine to form new two new orbitals called molecular orbitals. In the MOs one is lower in energy than the atomic orbitals and the other is higher in energy.

20. Fill in the electronic configuration for molecular oxygen (O_2) using the molecular orbital energy diagram given below.
$(\sigma_{2s})^2(\sigma_{2s}^*)^2(\sigma_{2p})^2(\pi_{2p})^2(\pi_{2p})^2(\pi_{2px}^*)^1(\pi_{2py}^*)^1$
b. What is the bond order for the molecule? **2**

c. Is the molecule paramagnetic or diamagnetic? Explain. Paramagnetic because there are 2 unpaired electrons in the MO diagram.

Chemistry 11, Section D01B, Spring 2006, Quiz #1

1. What is the chemical symbol for

copper ______ phosphorus_________

cobalt ______ chromium__________

2. Give the number of neutrons, protons, and electrons in the atoms of each of the following isotopes:

carbon-14

${}^{206}_{82}Pb$

3. Write the symbols for the ions of:

K_____ Br____ Mg___ S_____ Al____

4. Name the following compounds:

ClF_3________________________________

S_2Cl_2________________________________

Li_3N________________________________

K_2Se________________________________

Chemistry 11, Section D01B, Spring 2006, Quiz #1

1. What is the chemical symbol for

copper Cu phosphorus P

cobalt Co chromium Cr

2. Give the number of neutrons, protons, and electrons in the atoms of each of the following isotopes:

carbon-14 6 protons, 6 electrons, 8 neutrons

$^{206}_{82}Pb$ 82 protons, 82 electrons, 124 neutrons

3. Write the symbols for the ions of:

K K^{+} Br Br^{-} Mg Mg^{+2} S S^{-2} Al Al^{+3}

4. Name the following compounds:

ClF_3 chlorine trifluoride

S_2Cl_2 disulfur dichloride

Li_3N lithium nitride

K_2Se potassium selenide

Chemistry 11 Spring 2006 Quiz #2 Be sure to show your work to receive credit.

1. How many significant figures do the following numbers have?
a. 1.0230 kg   b. 3.0200 m   c. 0.04210 mm

2. Perform the following calculations and round the answers to the correct number of significant figures. Report the answers using the correct units.
a. (0.0023 m)(315 m)   b. 84.25 kg - 0.01075 kg

3. Report the following numbers in scientific notation.
a. 4350   b. 0.003287

4. Perform the following conversions. Tables of prefixes and conversions are given below.
a. 183 nm to cm   b. 3.1 ft. to meters.

Prefixes

pico	p	10^{-12}			
nano	n	10^{-9}	kilo	k	10^{3}
micro	µ	10^{-6}	mega	M	10^{6}
milli	m	10^{-3}	giga	G	10^{9}
centi	c	10^{-2}	tera	T	10^{12}

Conversions
1 inch = 2.54 cm
1 foot = 12 inches
1 yard = 3 feet
1 min = 60 sec
1 hr = 60 min

Chemistry 11 Spring 2006 Quiz #2 Be sure to show your work to receive credit.

1. How many significant figures do the following numbers have?

a. 1.0230 kg   b. 3.0200 m   c. 0.04210 mm

5   **5**   **4**

2. Perform the following calculations and round the answers to the correct number of significant figures. Report the answers using the correct units.

a. 0.0023 m x 315 m   b. 84.25 kg - 0.01075 kg

0.72 m^2   **84.24 kg**

3. Report the following numbers in scientific notation.

a. 4350   b. 0.003287

4.35 x 10^3   **3.287 x 10^{-3}**

4. Perform the following conversions. Tables of prefixes and conversions are given below.

a. 183 nm to cm   b. 3.1 ft. to meters.

183 nm(10^{-9}m/1 nm)(1 cm/0.01 m)
=**1.83 x 10^{-5} cm**

3.1 ft(12 in/1 ft)(2.54 cm/1 in)(0.01 m/1 cm)
= 0.94 m

Prefixes

pico	p	10^{-12}			
nano	n	10^{-9}	kilo	k	10^{3}
micro	μ	10^{-6}	mega	M	10^{6}
milli	m	10^{-3}	giga	G	10^{9}
centi	c	10^{-2}	tera	T	10^{12}

Conversions

1 inch = 2.54 cm
1 foot = 12 inches
1 yard = 3 feet
1 min = 60 sec
1 hr = 60 min

Chemistry 11 Spring 2006 Quiz #3

Be sure to show your work in detail when answering the following questions.

1. Calculate the % composition by mass for the following compound:
$CaSO_4$

2a. How many moles of sucrose ($C_{12}H_{22}O_{11}$) are in 150.00 g of sucrose?

b. How many molecules of sucrose are in 150.00 g of sucrose?

3. A compound is found to have a mass composition of 25.94% nitrogen and 74.06% oxygen. What is the empirical formula of this compound?

Chemistry 11 Spring 2006 Quiz #3

Be sure to show your work in detail when answering the following questions.

1. Calculate the % composition by mass for the following compound:
$CaSO_4$

The molar mass of CaSO4 = 40.08g/mol + 32.07g/mol + 4(16.00 g/mol) =

$$\%Ca = \frac{40.08 g/mol}{136.15 g/mol} x100 = 29.43\%$$

$$\%S = \frac{32.07 g/mol}{136.15 g/mol} x100 = 23.55\%$$

$$\%O = \frac{4(16.00 g/mol)}{136.15 g/mol} x100 = 47.01\%$$

2a. How many moles of sucrose ($C_{12}H_{22}O_{11}$) are in 150.00 g of sucrose?

$$150.00\, g\, C_{12}H_{22}O_{11} \frac{1\, mol\, C_{12}H_{22}O_{11}}{342.30\, g\, C_{12}H_{22}O_{11}} = 1.4355\, mol\, C_{12}H_{22}O_{11}$$

b. How many molecules of sucrose are in 150.00 g of sucrose?

$$1.4355 \text{ mol } C_{12}H_{22}O_{11} \frac{6.022 x 10^{23} molecules}{1 mol} = 8.645 \text{ x } 10^{23} \text{ molecules}$$

---

3. A compound is found to have a mass composition of 25.94% nitrogen and 74.06% oxygen. What is the empirical formula of this compound?

$Assume\, 100 g$

$$25.94 g\, N \frac{1 mol\, N}{14.01 g\, N} = 1.852\, mol\, N$$

$$74.06\, g\, O \frac{1 mol\, O}{15.999\, g\, O} = 4.629\, mol\, N$$

Divide by smallest #:

$$N: \frac{1.852 mol}{1.852 mol} = 1.00$$

$$O: \frac{4.629 mol}{1.852 mol} = 2.499 = 2.5$$

Multiply each value by 2 to eliminate the 0.5 for oxygen

N: 2(1.00) = 2 O: 2(2.50) = 5

Answer: N_2O_5

Chemistry 11
Spring 2006
Quiz #4

1. A sample of copper was heated to 120°C and then placed into 150 g of water at 20.00°C. The temperature of the mixture became 26.50°C.

a. How much heat in joules was absorbed by the water?

b. The copper sample lost how many joules?

c. What was the mass in grams of the copper sample?

specific heats

Substance	s (J/g°C)
carbon	0.711
copper	0.387
ethyl alcohol	2.45
gold	0.129
granite	0.803
iron	0.445
lead	0.128
olive oil	2.0
silver	0.235
water (liquid)	4.18

$q = ms\Delta T$

2. Given the following thermochemical equations:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H° = -571.5$ kJ
$N_2O_5(g) + H_2O(l) \rightarrow 2HNO_3(l)$ $\Delta H° = -76.6$ kJ kJ
$1/2N_2(g) + 3/2O_2(g) + 1/2H_2(g) \rightarrow HNO_3(l)$ $\Delta H° = -174.0$ kJ

calculate $\Delta H°$ for the reaction
$2N_2(g) + 5O_2(g) \text{-->} 2N_2O_5(g)$

General Chemistry 1
Quiz 4
1. A sample of copper was heated to 120.00°C and then placed into 150. g of water at 20.00°C. The temperature of the mixture became 26.50°C.

a. How much heat in joules was absorbed by the water?

$q = ms\Delta T = (150\ g)(4.18\ J/g°C)(26.50°C – 20.00°C) =$ **4080 J**

b. The copper sample lost how many joules?

$q_{Cu} = –q_{H2O} = –(4080\ J) =$ **–4080 J**

c. What was the mass in grams of the copper sample?

$m = q/s\Delta T = –4080\ J/(0.387\ J/g°C)(26.50°C – 120.00°C) =$ **113 g**

2.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$	$\Delta H° = – 571.5$ kJ
$N_2O_5(g) + H_2O(l) \rightarrow 2HNO_3(l)$	$\Delta H° = –76.6$ kJ kJ
$1/2N_2(g) + 3/2O_2(g) + 1/2H_2(g) \rightarrow HNO_3(l)$	$\Delta H° = –174$ kJ

calculate ΔH° for the reaction
$2N_2(g) + 5O_2(g) \rightarrow 2N_2O_5(g)$

1. Reverse the first equation and multiply by 2. This will give 2N2O5 as product.
2. Multiply the third equation by 4. This will give 2N2 as reactant.
3. Reverse the first equation. This will cancel the 2H2O.

$2[2HNO_3(l) \rightarrow N_2O_5(g) + H_2O(l)]$	$\Delta H° = –2(–76.6$ kJ kJ$) = +153.2$ kJ
$4[1/2N_2(g) + 3/2O_2(g) + 1/2H_2(g) \rightarrow HNO_3(l)]$	$\Delta H° = 4(–174$ kJ$) = –696$ kJ
$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$	$\Delta H° = –(– 571.5$ kJ$) = +571.5$ kJ

Total: 153.2 kJ – 696 kJ + 571.5 kJ **= +28.7 kJ**

Chemistry 11
Spring 2006
Quiz #5

1. How does the kinetic theory explain the existence of an absolute zero, 0 K? Offer an explanation.

2. When a sample of neon with a volume of 0.648 L and a pressure of 0.985 atm was heated from 15.0°C to 63.0°C, its volume changed to 0.755 L. What was its final pressure in atm? Be sure to show your reasoning.

$PV = nRT$

$R = 0.08206 \frac{\text{L atm}}{\text{K mol}}$

$T_K = T_{°C} + 273.15$

Chemistry 11
Spring 2006
Quiz #5

1. How does the kinetic theory explain the existence of an absolute zero, 0 K?
The kinetic theory states that the average kinetic energy of a collection of gas particles is related to (proportional to) the temperature of the particles. As the temperature of the gas particles decreases the average kinetic energy of the particles also decreases. Kinetic energy is proportional to velocity squared. So at some point the average velocity will approach zero and the average kinetic energy will also approach zero. The temperature at which this limit of zero velocity and zero kinetic energy occurs is called absolute zero (0 K).

2. When a sample of neon with a volume of 0.648 L and a pressure of 0.985 atm was heated from 15.0°C to 63.0°C, its volume changed to 0.755 L. What was its final pressure in atm?

This is a change of state problem. The volume changes. The pressure changes. The temperature changes. The number of moles of the gas are constant.

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

Because n is constant the equation simplifies to:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

We want P_2, the final pressure. Rearranging the equation to isolate P_2 gives:

$$P_2 = \frac{P_1V_1T_2}{V_2T_1} = \frac{(0.985\,\text{atm})(0.648\,\text{L})(336\,\text{K})}{(0.755\,\text{L})(288\,\text{K})} = 0.986\,\text{atm}$$

PV = nRT

$$R = 0.08206\frac{\text{L atm}}{\text{K mol}}$$

$T_K = T_{°C} + 273.15$

Chemistry 11
Spring 2006
Quiz #6

1. Describe the similarities and differences between **dipole-dipole interactions** and **London dispersion forces**.

2. What effect does raising the temperature of a liquid have on its equilibrium vapor pressure? Why?

Chemistry 11
Spring 2006
Quiz #6

1. Describe **dipole-dipole interactions** and **London dispersion forces**.

Dipole-dipole interactions are the attractive forces between molecules that possess a permanent dipole. London dispersion forces are the attractions between molecules (or atoms) that do not possess a permanent dipole. In London dispersion forces the attractions are caused by spatial fluctuations in the electronic distribution of the electrons. These fluctuations can result in dipoles that can induce dipoles in other particles. the net result is induced dipoles in the materials and attractions.

2. What effect does raising the temperature of a liquid have on its equilibrium vapor pressure? Why?

An increase in the temperature increases the vapor pressure of a liquid. This is because increasing the temperature increases the average kinetic energy of the molecules (or atoms) in the liquid. With an increased average kinetic energy a larger proportion of the particles in the liquid possess enough kinetic energy to escape the surface of the liquid and move into the gaseous state.