CHEMISTRY PILOT REVIEW SHEET SECOND SEMESTER

Utica Community Schools – Semester Two Review

Directions:

- You will be given a periodic table with polyatomic ions and any mathematical formulas you may need.
- If you complete the entire review sheet by the day of your exam you will be eligible for the curve should one be applied.
- The answers to this review must be on a **separate sheet of lined paper**. To the best of our ability we have tried to provide a chapter that can be reviewed to help with that content.

Chapter 2: Antacids

1. Know how to identify a compound as an ionic or covalent compound. Covalent compounds are also referred to as molecular compounds.
    a. Circle the following formulas that represent ionic compounds. Put a box around those that are covalent/molecular.

    NO_2 NaCl $BaCl_2$ CH_4 $PbSO_4$ H_2O_2 Fe_2O_3 $(NH_4)_2CO_3$ PH_3

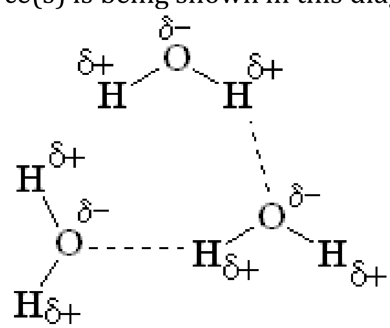
2. Know how to write the names and formulas for ionic compounds, covalent compounds, acids and bases.
    a. Write the names for the following compounds:
        i. S_4O_6 iii. KOH v. $Mg(CN)_2$
        ii. $Fe_3(PO_4)_2$ iv. CO
    b. Write the formulas for the following compounds:
        i. Silicon dioxide iii. potassium fluoride v. cesium hypochlorite
        ii. Lithium sulfite iv. Tricarbon octahydride vi. Iron (II) oxide
    c. Why do certain compounds require a roman numeral? What block of elements requires them?

Chapter 5: Soap

3. For each example below, predict if the bonding will be ionic or covalent.
    i. Ca with Cl
    ii. C with H
    iii. Carbon monoxide
    iv. KBr
    v. SO_2
    vi. Magnesium iodide
4. Draw Lewis dot structures for these simple compounds listed below.
    i. Br_2
    ii. O_2
    iii. N_2
    iv. CH_4
    v. NH_3
    vi. H_2O
    vii. HF

5. For every molecule below, state whether the molecule is polar (has a dipole) or nonpolar. Look for symmetry – a sign of being NONPOLAR! A lone pair on the central atom generally leads to polarity or having a dipole.
    i. Cl_2
    ii. O_2
    iii. CH_4
    iv. NH_3
    v. H_2O
    vi. HF

6. Explain why oxygen is gas but water is a liquid at room temperature.
7. Which state of matter has the strongest intermolecular forces of attraction? Which state has the weakest? Why is water a liquid at room temperature but carbon dioxide is a gas?
8. Explain why fluorine and chlorine are gases but bromine is a liquid and iodine is a solid.
9. What three elements can hydrogen bond with to create the IMF of Hydrogen Bonding BETWEEN molecules?
10. What type of intermolecular force(s) is being shown in this diagram?



11. The following questions all pertain to intermolecular forces. Please answer each question.
    a. What is a London Dispersion force?
    b. What is a Dipole-Dipole Attraction?
    c. What is Hydrogen Bonding?
    d. Rank the three intermolecular forces from weakest to strongest.
    e. What types of intermolecular forces are present on each of the following molecules: You will need the Lewis Dot Drawings and Polarity information from questions 3 and 4.
        i. Cl_2
        ii. O_2
        iii. CH_4
        iv. NH_3
        v. H_2O
        vi. HF

12. What type of compounds (IONIC OR COVALENT) are sugar ($C_6H_{12}O_6$), ammonia (NH_3), water (H_2O) and butane (C_4H_{10})? Why is sugar a solid at room temperature yet ammonia is a gas? Why does water boil at a higher temperature than butane?
13. What happens to the strength of the intermolecular forces as water is freezing? As water is melting?
14. True or False: Solids have greater intermolecular forces than liquids and gases.
15. True or False: As you go down the halogen family, intermolecular forces increase.
16. Fill in the blank: The higher the intermolecular forces the __________________ the melting point and the __________________the boiling point.
17. True or False: The IMF with the highest boiling points will have hydrogen bonding.
18. Rank these three molecules from highest boiling point to lowest: HCl, CH_4, H_2O. Why did you choose this order?

19. The molecular formula and boiling points of hydrocarbons are given in the table below.

Table 1 *Hydrocarbon Boiling Points*

Hydrocarbon	**Formula**	**Boiling Point (°C)**
Butane	C_4H_{10}	-0.5
Decane	$C_{10}H_{22}$	174.0
Ethane	C_2H_6	-88.6
Heptane	C_7H_{16}	98.4
Hexane	C_6H_{14}	68.7
Methane	CH_4	-161.7
Nonane	C_9H_{20}	150.8
Octane	C_8H_{18}	125.7
Pentane	C_5H_{12}	36.1
Propane	C_3H_8	-42.1

Based on the information in the table, what is the trend in the boiling point as the number of carbons increase? Based on IMFs, why is this so?

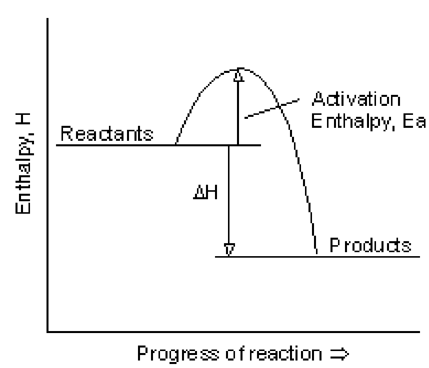
<u>Chapter 7: Hot Packs</u>

20. For the reaction given below, list the bonds that are broken. Make a list of the bonds being formed:

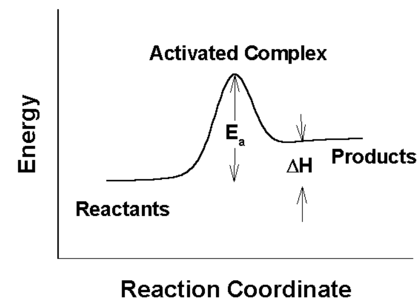
$$Ba(OH)_2 + 2NH_4Cl \rightarrow BaCl_2 + 2H_2O + 2NH_3$$

21. List single, double, and triple bonds in order of increasing strength. In order of increasing length.
22. Describe the bond between N_2 using these characteristics.
23. Label the following phase changes as exothermic or endothermic.
    i. Water is boiling, energy is gained by the system. ΔH is positive
    ii. Water is condensing, energy is lost to the surroundings. ΔH is negative
    iii. Water is freezing, energy is lost to the surroundings. ΔH is negative.

24. Are these reactions exothermic or endothermic?
    i. $2H_2 + O_2 \rightarrow 2H_2O + 483.6\ kJ$
    ii. $N_2 + 2O_2 + 34kJ \rightarrow 2NO_2$
    iii. Heat + $H_2O_2 \rightarrow H_2O + O_2$
    iv. $S + O_2 \rightarrow SO_2$ + Heat
25. Is the diagram shown below exothermic or endothermic?
    a. For an exothermic reaction are the reactants higher or lower than the products?
    b. What would happen to the temperature of this reaction mixture?



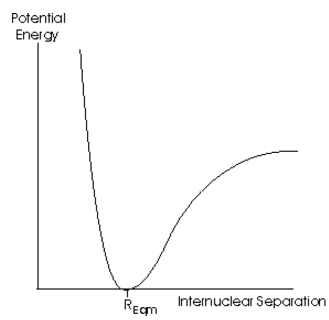
26. Is the diagram shown below endothermic or exothermic?
    a. Would the ΔH be positive or negative?
    b. In an endothermic reaction are the reactants higher or lower than the products.



27. Explain why freezing is an exothermic change of state.

28. Study the diagram below. Answer the questions that follow:

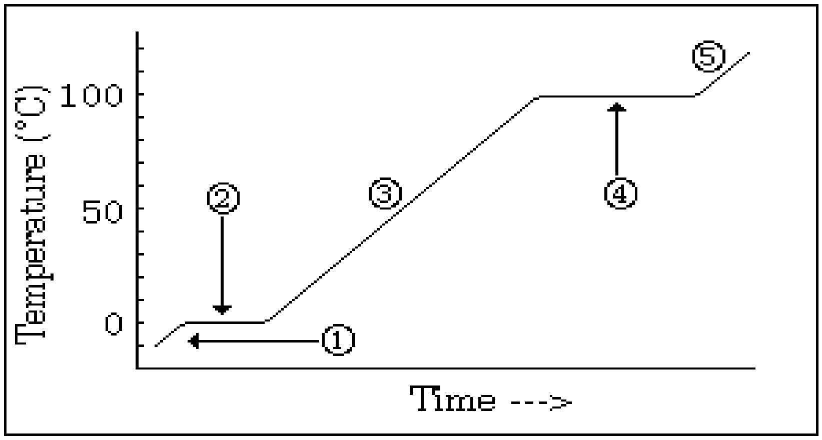
    Potential Energy Diagram in the formation of a Hydrogen Molecule



Point R, the lowest point on this diagram, represents the distance at which a bond has reached its lowest potential energy and hence represents the distance at which the bond will form. At R, the bond is formed but notice that to the left and right of the R, the potential energy rises. This means that as bonds are pulled apart, the potential energy rises. As the bond is pushed to close together, the potential rises.

i. Where is the bond most stable? Before R, at R or after R?
ii. What happens to potential energy as we move away from point R
iii. When atoms bond, they achieve a lower, more stable energy – True or False?
iv. As atoms are pulled apart from bonding, their potential energy increases. True or False?

29. Answer the questions below about the following temperature versus time graph of water.



i. What state(s) of matter are shown at all numbered regions on this graph?
ii. Identify all regions on the graph where average kinetic energy is increasing.
iii. Identify all regions on the graph where only potential energy is increasing.
iv. During phase 2, as energy is being added, what is happening to the substance? Why is this process endothermic?
v. When ice melts, the kinetic energy of the water molecules must break what types of forces?
vi. What instrument could you use to make the measurements on this curve?

NGSS

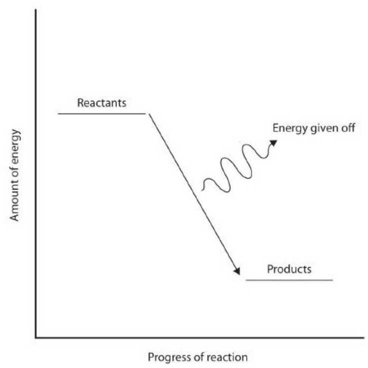
30. What is the definition of a chemical reaction?
31. What evidence can we observe that suggests a chemical reaction has occurred?
32. What is the definition of a physical change?
33. What happens to a solid when it dissolves? Is this a physical or chemical change?
34. Define Kinetic Energy:
35. Define Temperature:
36. Define Energy:
37. Define Thermal Energy:
38. Define endothermic:
39. Define exothermic
40. Why do magnets serve as a model for the physical connection between atoms?
41. Is it necessary to have more energy if we want to dissolve more?
42. Complete the following sentences using the following word bank: Hotter, colder, faster, slower, more easily, less easily

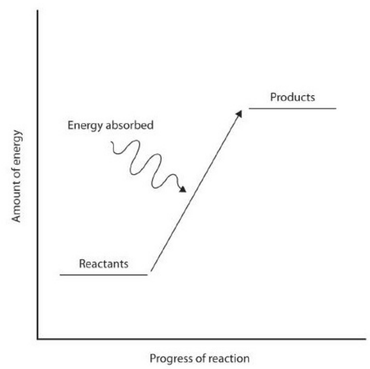
    The__________________ the water, the _________________ the water particles, the _______________________ the substance is dissolved.

    The__________________ the water, the _________________ the water particles, the _______________________ the substance is dissolved.

43. Which process **releases** energy: breaking a bond or forming a bond?
44. Which process **requires** energy: breaking a bond or forming a bond?

45. What are some properties of alkali metals?
    a. What gas was produced when alkali metals were added to water?
    b. What type of reaction was this?
46. Based on the two pictures below, state which one is endothermic and which exothermic?
    a. How you can tell?





47. Is each chemical reaction exothermic or endothermic?

a. $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)} + 135\ \text{kcal}$

b. $N_{2(g)} + O_{2(g)} + 45\ \text{kcal} \rightarrow 2NO_{(g)}$

Chapter 8: Chemistry in Industry or Equilibrium

48. Use LeChatelier's Principle to answer the following set of questions:

$$2NO_{(g)} + O_{2(g)} \leftrightarrow 2NO_{2(g)} + \text{heat}$$

    a. For the reaction shown above, what would happen to the concentration of reactants if pressure were increased on the system?
    b. If the NO_2 is increased, how will the equilibrium be affected? What will happen to the concentration of the oxygen gas?
    c. If more heat is added by raising the temperature, which way will the reaction shift?
    d. Is the above reaction, endo or exo-thermic?

Chapter 3: Airbags

49. Know how to perform mole conversions using moles, grams, and representative particles.
    a. What is a representative particle? What is the name of the representative particle for the ionic compound? The covalent compound? An element?
    b. How many representative particles are in a 53.79 gram sample of $SrCO_3$?
    c. How many moles of AlF_3 are in a 10.0 gram sample of AlF_3?
    d. How many grams of SO_2 are in a 0.65 mole sample of SO_2?
    e. How many moles of copper atoms in 75 g of Cu?
    f. In a sample with 17.0 grams of water, how many molecules of water are present?

Chapter 6: Sports Drinks

50. Use the following balanced equation to answer the following problems:

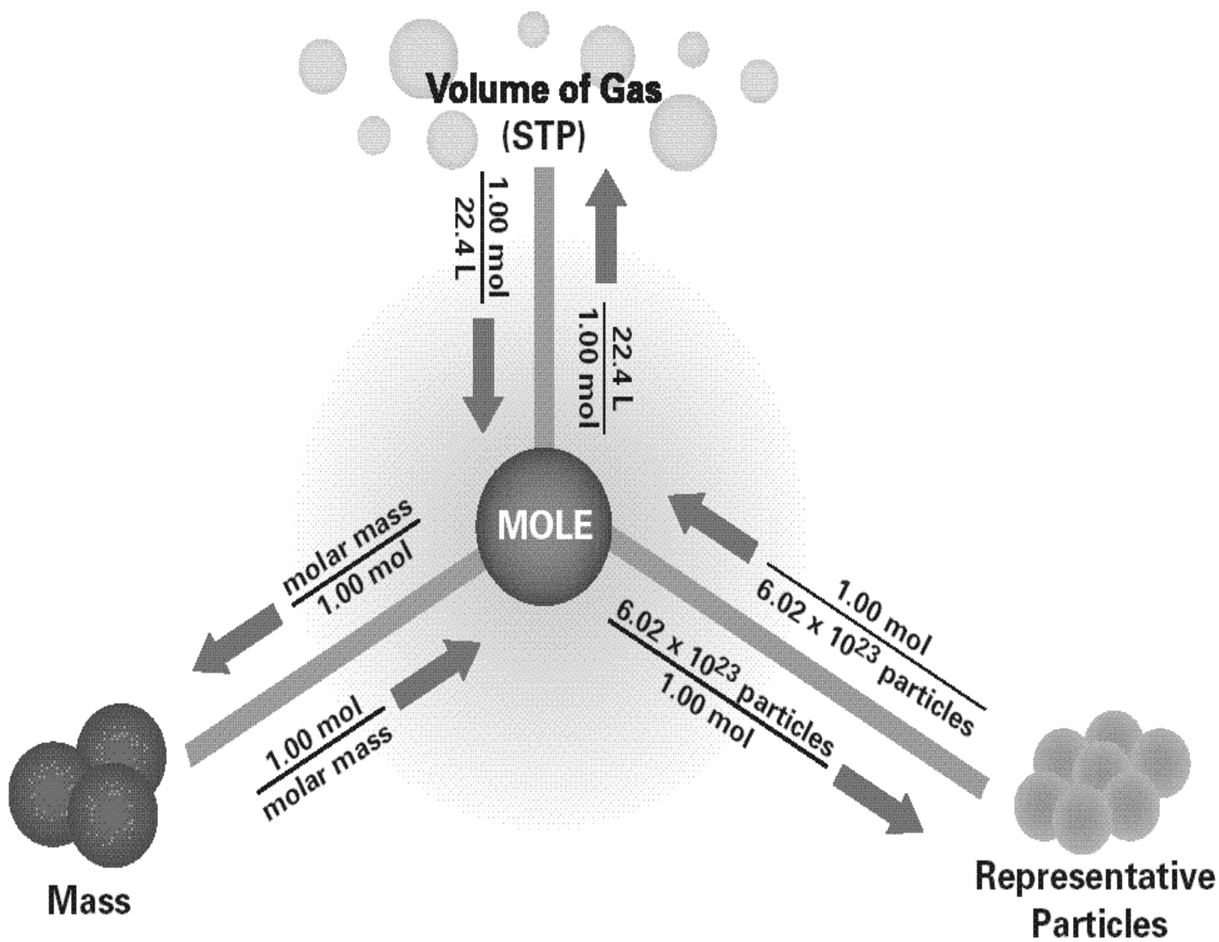
$$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$$

i) When 2.5 moles of oxygen gas react completely with pentane, how many moles of carbon dioxide are produced?

ii) How many grams of pentane gas are needed to completely react with 348.5 grams of oxygen?

iii) How many L of CO_2 are produced if you start with 6.75×10^{26} molecules of oxygen?

1 mol = 22.4 L   1 mol = 6.02×10^{23} representative particles   1 mol = g (molar mass)



Periodic Table of the Elements

Chemistry Reference Sheet

California Standards Test

Key: 11 — Atomic number; Na — Element symbol; Sodium — Element name; 22.99 — Average atomic mass*

	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 H Hydrogen 1.01																	2 He Helium 4.00
2	3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
3	11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
4	19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
5	37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
6	55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)									

58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97
90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	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)

* If this number is in parentheses, then it refers to the atomic mass of the most stable isotope.

POLYATOMIC IONS:

Ammonium	NH_4^{+1}		Cyanide	CN^{-1}
Acetate	$C_2H_3O_2^{-1}$	CH_3COO^{-1}	Phosphate	PO_4^{-3}
			Phosphite	PO_3^{-3}
Bicarbonate	HCO_3^{-1}		Hydroxide	OH^{-1}
Carbonate	CO_3^{-2}		Nitrate	NO_3^{-1}
Perchlorate	ClO_4^{-1}		Nitrite	NO_2^{-1}
Chlorate	ClO_3^{-1}		Sulfate	SO_4^{-2}
Chlorite	ClO_2^{-1}		Sulfite	SO_3^{-2}
Hypochlorite	ClO^{-1}		Bisulfate	HSO_4^{-1}

Covalent Naming Prefixes:

1- Mono
2- Di
3- Tri
4- Tetra
5- Penta
6- Hexa
7- Hepta
8- Octa
9- Nona
10- deca

Ionic Naming

Metals in column 1, 2 and Al, Zn and Ag – no Roman numerals needed!

Rest of metals will need a Roman numeral in ( ) to indicate its charge.