CHEMISTRY 2016-2017 FINAL EXAM TOPICS

OUTLINE OF TOPICS:

1. Reading chemical equations (formulas, states of matter, etc)
2. Stoichiometry
    a. dimensional analysis (or the factor-label method)
    b. Theoretical yield
3. Lewis dot structures
    a. Molecular polarity
4. Electronegativity
    a. Bond polarity
5. States of matter
    a. entropy
    b. Phase changes
    c. Phase change diagrams
6. Kinetics
    a. reaction rates and factors affecting reaction rate
    b. open vs closed system
    c. Endothermic reactions
    d. Exothermic reactions
    e. Activation energy
    f. Catalyst
7. Hydrogen bonding
8. Solubility
    a. Factors that affect solubility
9. pH scale of acids and bases
10. Concentration of solutions
    a. Molarity
    b. Dilution calculations
11. Methods for separating mixtures
12. Behavior of gases as V,T, P, or concentration changes
    a. Gas law calculations
    b. STP
    c. Unit conversions

LEARNING STANDARDS:

HS-PS1-7. Use mathematical representations and provide experimental evidence to support the claim that atoms, and therefore mass, are conserved during a chemical reaction. Use the mole concept and proportional relationships to evaluate the quantities (masses or moles) of specific reactants needed in order to obtain a specific amount of product.

- Evaluations may involve mass-to-mass stoichiometry.

HS-PS1-4. Develop a model to illustrate the energy transferred during an exothermic or endothermic chemical reaction based on the bond energy difference between bonds broken (absorption of energy) and bonds formed (release of energy).

- Examples of models may include molecular-level drawings and diagrams of reactions or graphs showing the relative energies of reactants and products.

HS-PS1-5. Construct an explanation based on kinetic molecular theory for why varying conditions influence the rate of a chemical reaction or a dissolving process. Design and test ways to slow down or accelerate rates of processes (chemical reactions or dissolving) by altering various conditions.*

- Explanations should be based on three variables in collision theory: (a) quantity of collisions per unit time, (b) molecular orientation on collision, and (c) energy input needed to induce atomic rearrangements.
- Conditions that affect these three variables include temperature, pressure, and concentrations of reactants, agitation, particle size, surface area, and addition of a catalyst.

HS-PS3-4b. Provide evidence from informational text or available data to illustrate that the transfer of energy during a chemical reaction in a closed system involves changes in energy dispersal (enthalpy change) and heat content (entropy change) while assuming the overall energy in the system is conserved.

HS-PS1-3. Cite evidence to relate physical properties of substances at the bulk scale to spatial arrangements, movement, and strength of electrostatic forces among ions, small molecules, or regions of large molecules in the substances. Make arguments to account for how compositional and structural differences in molecules result in different types of intermolecular or intramolecular interactions.

- Types of intermolecular interactions include dipole-dipole (including hydrogen bonding), ion-dipole, and dispersion forces.

HS-PS2-8(MA). Use kinetic molecular theory to compare the strengths of electrostatic forces and the prevalence of interactions that occur between molecules in solids, liquids, and gases. Use the combined gas law to determine changes in pressure, volume, and temperature in gases.

HS-PS1-9(MA). Relate the strength of an aqueous acidic or basic solution to the extent of an acid or base reacting with water as measured by the hydronium ion concentration (pH) of the solution. Make arguments about the relative strengths of two acids or bases with similar structure and composition.

Clarification Statements:

- Reactions are limited to Arrhenius and Bronsted-Lowry acid-base reaction patterns with monoprotic acids.
- Comparisons of relative strengths of aqueous acid or base solutions made from similar acid or base substances is limited to arguments based on periodic properties of elements, the electronegativity model of electron distribution, empirical dipole moments, and molecular geometry. Acid or base strength comparisons are limited to homologous series and should include dilution and evaporation of water.

HS-PS1-11(MA). Design strategies to identify and separate the components of a mixture based on relevant chemical and physical properties.

HS-PS2-7(MA). Construct a model to explain how ions dissolve in polar solvents (particularly water). Analyze and compare solubility and conductivity data to determine the extent to which different ionic species dissolve.

1. Calculations bases on quantitative relationships and balanced chemical equations are called ***stoichiometry***.
Stoichiometric calculations are bases on the **mole ratios** between any two substances in the reaction.
What is the mole ratio between oxygen and water in the following reaction? $2H_2 + O_2 \rightarrow 2H_2O$

2. Use the **mole *map*** to help you solve conversions (don't forget mole ratios!) between moles, grams, numbers of molecules/atoms, and liters of gases at STP.
Given the reaction $CH_4 + 2O_2 \dashrightarrow CO_2 + 2H_2O$,
what amount of carbon dioxide is produced by the reaction of 1 mole of CH_4*?*
1 gram *1 liter* *1 mole* *22 grams*

3. ***Reactants*** are on the left side of the reaction arrow and ***products*** are on the right.
What are the names of the reactants in this neutralization reaction?
$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

4. ***Lewis Electron dot structures*** represent covalently bonded molecules formed through the sharing of valence electrons between atoms. It also shows the lone-pair(s) of electrons that may affect the shape and polarity of the molecules.
Draw the Lewis structures for these common molecules H_2 H_2O NH_3 CH_4

5. ***Exothermic reactions*** release energy (energy is a product of the reaction) while ***Endothermic reactions*** absorb energy and the *energy is a reactant* in the reaction.
Given the reaction: $CH_4\,(g) + 2\,O_2\,(g) \rightarrow 2\,H_2O\,(g) + CO_2\,(g) + \text{heat}$
What is the overall result when $CH_4\,(g)$ *burns according to this reaction?*
Energy is absorbed *Energy is released*

6. The gram formula mass (**molar mass**) of a substance is the sum of the atomic masses of all the atoms in it.
H_2SO_4 = ______ g/mole

7. ***6.02 x*** 10^{23} is called ***Avogadro's number*** and is the number of particles in ***1 mole*** of a substance.
Equal volumes of gases contain an equal number of molecules.
Under similar conditions, which sample contains the same number of moles of particles as 1 liter of $O_2\,(g)$*?* *1 L* $Ne(g)$ *0.5 L* $SO_2\,(g)$ *2 L* $N_2\,(g)$ *1 L* $H_2O(l)$

8. The **kinetic molecular theory** explains the behavior of matter as particles with energy and motion.

9. ***Temperature*** is a measure of average kinetic. *Which sample has the highest average kinetic energy?*
$H_2O\,(l)$ *at 0°C* $H_2O\,(s)$ *at 0°C* $CO_2\,(g)$ *at STP* *Mg (s) at 298K*

10. ***Reaction rate*** increases with an increase in temperature (and pressure for gases).

11. ***Catalysts*** speed up reactions by lowering their ***activation energies***.
They are not changed themselves and can be reused many times over.

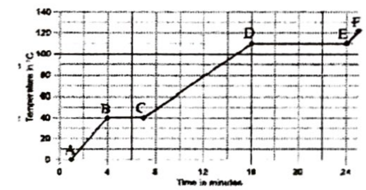
12. The particles in a ***solid*** are rigidly held together, closely packed in a **lattice** arrangement.
Which of the following has a regular geometric arrangement at 298 K and 1.0 atm?
$Br_2\,(l)$ $CO_2\,(g)$ *Mg (s)* $H_2O\,(l)$

13. ***Solids*** have a definite shape and volume.

14. ***Liquids*** have closely-spaced particles that easily slide past one another; they have no definite shape, but have a definite volume.

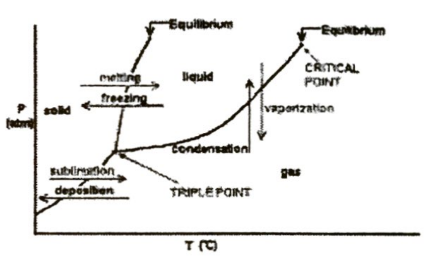
15. ***Gases*** have widely-spaced particles that are in random motion (collide with container to create pressure). ***Gases*** are easily compressed and have no definite shape or volume.

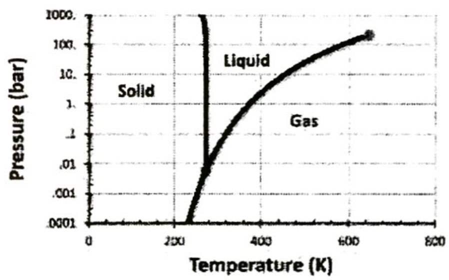
16. Be able to read and interpret heating/cooling curves as pictured below.



*During which **interval** on the graph are solid and liquid in equilibrium?*

17. Be able to read and interpret phase diagrams. At STP what phase of matter is the substance? At 1 atm, what is the melting point?





18. Substances that ***sublime*** turn from a solid directly into a gas. They have very weak attractive or intermolecular forces. (examples include CO_2 & I_2)
What type of intermolecular forces exist between molecules of CO_2 or I_2?

19. "***STP***" means "***S****tandard* ***T****emperature and* ***P****ressure.*"

These conditions define STP P = ______atm = _______kPa T = ________K

20. Degrees Kelvin = C + 273 ***Temperature*** is a measure of the **kinetic energy** of the particles in matter.
Room temperature = 25°C = ...K *Boiling point of helium = 4 K = ...°C*

21. ***Always use Kelvins*** for temperature when using the ***combined gas law.*** $P_1V_1T_2 = P_2V_2T_1$
Set up the equation to calculate the volume of 50. mL of methane gas collected at STP when the pressure rises to 2.4 atm and the temperature drops to 240 K.

22. As the ***pressure*** exerted on a gas increases, the ***volume*** decreases proportionally.
25 L of a gas is held at 1.2 atm pressure.
Find the new volume if pressure drops to 0.80 atm at constant temperature.

23. As the ***pressure*** on a gas increases, ***temperature*** increases.
A sample of gas exerts a pressure of 220. kPa at 373 K. Find the pressure at 273 K at constant volume.

24. As the ***temperature*** of a gas increAses, ***volume*** increases.
15 mL of oxygen gas is collected at 0°C. Find the volume at 50°C at constant pressure.

25. ***Real gas*** particles have volume and are attracted to one another. They don"t always behave like ***ideal gases.*** Lighter gases (with weaker attractive forces) are often most ideal.

Which of the following is the most ideal gas?

He *Ne* *Ar* *Kr*

26. Real gases behave more like ideal gases at *low pressures and high temperatures.*

27. According to ***Avogadro's law*** at constant temperature and pressure, the volume of a gas is directly proportional to the number of moles. This is true for any gas.

At STP 22.4 L of any gas = 1 mole, what is this equality called?

28. The ***Ideal Gas Law*** relates the pressure (atm), volume (L), temperature (K), and amount of gas particles (moles) of a gas. The formula is PV = nRT where R is 0.0821 L•atm/mol•K

What is the volume occupied by 36.0g of water vapor at 125°C and 0.999atm?

29. The last digit of an element's group number is equal to its ***number of valence electrons.***

Which contains the greatest number of valence electrons?

Ca *Ge* *Se* *Kr*

30. Draw one dot for each valence electron when drawing an element's or ion's ***Lewis electron dot diagram.***

Which dot model would contain the fewest dots as valence electrons?

Ca *Ge* *Se* *Kr*

31. ***Nonpolar covalent bonds*** form when two atoms of the *same element* bond together.

32. ***Polar covalent bonds*** form when the electronegativity difference between two bonding atoms is between 0.4 and 1.7. *Which of the following combinations would form a polar covalent bond?*

H and H *Na and N* *H and N* *Na and Br*

33. ***Ionic bonds*** form when the electronegativity difference between two bonding atoms is *greater than* 1.7.

Which of the following combinations would form an ionic bond?

H and H *Na and N* *H and N* *Na and Br*

34. **Polar molecules** have stronger forces of attraction. They lack structural symmetry.

Which of the following is a polar molecule?

CO_2 H_2O C_4H_{10} N_2

35. ***Hydrogen bonds*** are attractive forces that form when hydrogen bonds to the elements N, O, or F and gives the compound unexpectedly high melting and boiling points.

The strongest forces of attraction occur between molecules of

HCl *HBr* *HF* *HI*

36. ***Solutions*** are the best examples of ***homogeneous mixtures.*** They have **two** sets of properties.

37. ***Heterogeneous mixtures*** have discernible components and ***are not*** uniform throughout.

Air is classified chemically as a(n)

Substance *compound* *element* *solution*

38. A ***solute*** is the substance being dissolved; the ***solvent*** is the substance that dissolves the solute.
NaCl (s) is added to water.
The solute is ..... *the solvent is ......* *the solution is ....*

39. Remember: substances tend to be soluble in solvents with similar molecular properties.
"Like dissolves like"
Pentane does not dissolve in water because pentane is .... and water is ....

40. As temperature increases, solubility increases for most solids.
For which substance does increasing temperature have the least effect on solubility?
Sodium chloride *calcium carbonate* *oxygen* *sodium bicarbonate*

41. At low temperatures and high pressures solubility *increases* for most gases.
*Carbon dioxide gas is **least** soluble in water at conditions of .... temperature and .... pressure.*

42. ***Molarity*** is a way to measure the *concentration* of a solution.
Molarity is equal to the number of moles of solute divided by the number of liters of solution.
What is the molarity of an NaCl solution if 2.0 mol NaCl is present in 0.50 L solution?

43. ***Molality*** = (moles of solute/ kg of sovent)
A solution of glucose is prepared by added 10. g glucose ($C_6H_{12}O_6$) to 40. g water.
What is its molality?

44. ***Dilutions*** → $M_1V_1 = M_2V_2$ is the way to determine how to make a less concentrated solution.
A student needs 250 mL of a 1.0 M HCl solution, how many mL of 6.0M HCl do they need to make this solution?

45. Increasing the **concentration** of reactants will increase ***reaction rate***.
Which sample of HCl (aq) will react most rapidly with magnesium metal?
0.50 M HCl *1.0 M HCl* *3.0 M HCl* *6.0 M HCl*

46. ***Acids*** and ***bases*** are both ***good electrolytes***. Their solutions conduct electricity well.
Which of the following is a nonelectrolyte?
LiOH *HBr* $HC_2H_3O_2$ C_2H_5OH

47. **Weak acids** taste *sour* and *react with metals.*
48. **Weak bases** taste *bitter* and *feel slippery.*

49. **Acids** have a pH < 7. **Bases** have a pH > 7.

50. ***Brønsted*** model of acids and bases states:
"Acids *donate* protons." "Bases *accept* protons."
Identify one Bronsted acid and one Bronsted base in the reaction below:
$H_2O + NH_3 \Leftrightarrow NH_4^+ + OH$

Stoichiometry problems

1. Methanol, CH3OH, can be produced by the following reaction: $2H_2 + CO \dashrightarrow CH_3OH$

a) Calculate the theoretical yield of CH_3OH if 68.5 g of CO is reacted with 8.6 g of H_2. (2 givens and 2 calculations)

Theoretical yield = ______________

b) What is the limiting reactant in the reaction? The reactant in excess?

__________________ **is LR,** __________________ **is in excess**

c) If 35.7 g CH_3OH is actually produced, what is the % yield of methanol?

% Yield = $\dfrac{\textit{Actual Yield}}{\textit{Theoretical Yield}}$ **x 100% =**

2. ______$BaCl_2$ + ______Na_3PO_4 → ______$Ba_3(PO_4)_2$ + ______NaCl

a. Balance the equation above.
b. How many molecules of NaCl are produced when 3.98 mol of $BaCl_2$ reacts?

c. If 5.17×10^{30} molecules of Na_3PO_4 react, how many grams of $Ba_3(PO_4)_2$ are made?

d. If 10.9 moles of NaCl are produced in a reaction, how many moles of Na_3PO_4 were reacted?

Ideal Gas Problems

Gases at low pressures obey the ideal gas law,

$$p\,V = n\,R\,T \qquad (1)$$

where R is a constant (known as the *gas constant*) that has the value

$$R = 0.08206 \text{ L atm K}^{-1}\text{ mol}^{-1} \qquad (2)$$

Appropriate units to use for p, V, n, and T in the ideal gas equation are those used for R above. Thus the pressure (p) should be in atm, the volume (V) in L, the temperature (T) in degrees K, and the amount of gas (n) should be in moles. Useful conversion factors are

Pressure:	1 atm = 760 Torr = 760 mmHg = 101.3 kPa = 1.013 bar
Temperature:	K = 273 + ^{o}C
Volume:	1 L = 1000 mL = 1000 cm^3

STP

Often you will see gas volumes reported at STP (*standard temperature and pressure*). STP is defined as T = 273 K (0^{o}C) and p = 1 atm. Substitution of these values into Eq(1) shows that the *volume of 1 mol of any gas is approximately 22.4 L at STP*. (You should verify this for yourself using Eq(1)!).

1. A Marshmallow Peep® has a volume of about 45.0 cm^3 at 101 kPa. What pressure is required to increase its size to 150.0 cm^3 assuming no air escapes from the Peep®

2. What is the temperature of a 0.00893 mol sample of neon gas that has a volume of 302 mL and a pressure of 0.941 atm?

3. A gas occupies 4.78 L at 78.1 kPa and 25°C. What will the volume be at 0.975 atm and 15°C?

4. A shampoo bottle contains 443 mL of air at 65°C. What is its volume when it cools to 22°C?

5. The pressure in a can of hairspray is 2.50 atm at 298 K. What is the pressure in the can when it is heated to 398 K?

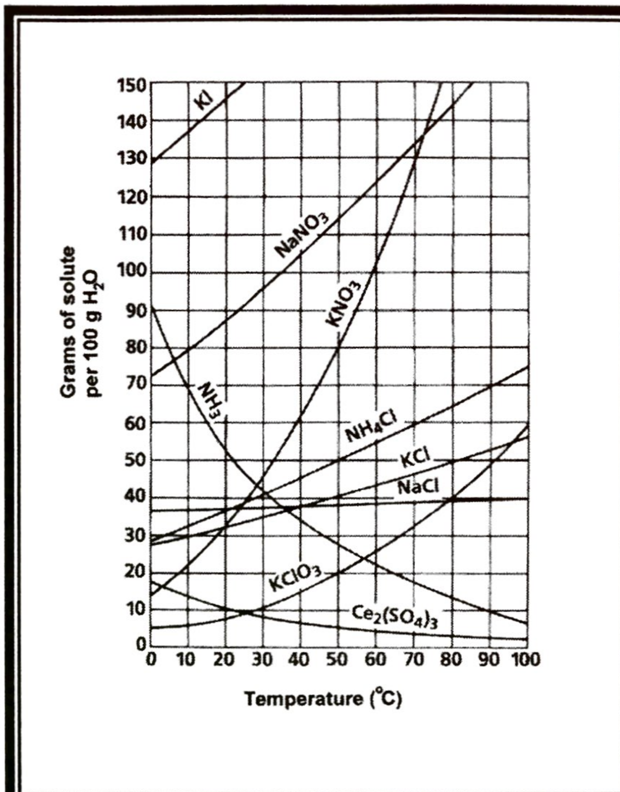
6. What mass of glucose ($C_6H_{12}O_6$) is required to produce 150 mL of carbon dioxide at 102 kPa and 23°C? How many molecules of glucose is this?

$$C_6H_{12}O_6 + 2O_2 \rightarrow 2CH_3COOH + 2CO_2 + 2H_2O$$

Name________________________________	Date ________________	Period ______

Review for Solutions, Acids, and Bases

Solutions



1. Explain the effect of adding more solute to unsaturated, saturated, and supersaturated solutions.
2. Explain how temperature and pressure affect solubility.

State whether each pair is soluble or insoluble.

3. KCl in water
4. ammonia in oil
5. wax in C_6H_6
6. CH_4 in water
7. Read solubility curves (See *the solubility curve*).
   a) What is the solubility of ammonium chloride at 60°*C*?
   b) At what temperature do potassium chlorate and potassium chloride have the same solubility in water?
   c) Which compound is least soluble in water at 12°*C*?
   d) A saturated solution of which compound contains 130 grams of solute per 100 grams of water at 70°*C*?
   e) Are the following solutions unsaturated, saturated, or supersaturated?
      1. 80 g of sodium nitrate in 100 g of water at 30°C.
      2. 80 g of potassium chlorate in 100 g of water at 50°C.
   f) How many grams of sodium chloride are required to saturate 500 grams of water at 100° *C*?
8. How many grams of $AlCl_3$ are required to make a 2.25m solution in 30.0 g of water?
9. What volume of 12*M* HCl is needed to prepare 250 mL of 0.20*M* HCl?
10. Explain the difference in preparing solutions based on molarity versus molality.
11. Which will have the greatest effect on Δt_f at the same molality: $C_{12}H_{22}O_{11}$, $MgBr_2$, $AlCl_3$, or NH_4NO_3?
12. When 26.4 g of NaBr dissolves in 0.20 kg of water, what is the freezing point of the solution?

VOCAB: **solvation** **solubility** **dissociation** **ionization** **molality** **strong/weak/nonelectrolyte** **molarity**

Acids and Bases

State whether the following are acids or bases.

13. Have a sour taste.
14. React with metals.
15. Feel slippery
16. Turn blue litmus paper red.
17. Define acids and bases according to Arrhenius, Brønsted-Lowry, and Lewis.
18. Identify each substance as acid, base, conjugate acid, or conjugate base.
    $H_2S + H_2O \rightarrow HS^- + H_3O^+$
19. Give the conjugate acids of: NH_3 and Br^-.
20. Give the conjugate bases of: H_3O^+ and HSO_4^-.
21. Find the pH of 0.75*M* HCl.
22. Find the molarity of a KOH solution with a pH of 9.5.
23. Is the solution in #22 acidic or basic?
24. What is the product in a neutralization reaction between HCl (aq) and NaOH (aq)?
25. If 43.5 mL of 0.15 M HBr is required to neutralize 25.0 mL of $Ca(OH)_2$, what is the molarity of $Ca(OH)_2$? *Skip 2014*

VOCAB: **hydronium ion**
neutralization reaction
amphoteric substance
titration
strong/weak acid/base
equivalence point