Accelerated Semester 2 Review **Name:** Key **Per.** ____

The performance task will be taken in class on Wednesday, June 1. The 62 question multiple choice test will be taken on Thursday, June 2nd for periods 4-7 and Friday, June 3rd for periods 1-3.

A. Unit ~~I~~-Acids and Bases

Vocabulary: Review the following vocabulary. Look up and write the definition for any words you do not know.

Arrhenius model	Bronsted-Lowry model	Conjugate base	Kw	pH
Acid-base indicator	Conjugate acid	End point	hydronium ion	pOH
Amphoteric (amphiprotic)	Conjugate acid-base pair	Equivalence point	neutralization reaction	titration

Unit Objectives:

- Distinguish between acids and bases as defined by Arrhenius and Bronsted-Lowry
- Distinguish between strong and weak acids and bases
- Explain the concept of neutralization & discuss how titrations can be used with acids and bases in neutralization reactions
- Explain and calculate pH and pOH (using H^+ and OH^- concentrations)
- Explain how buffers resist changes in pH

1. List 5 properties of acids and 5 properties of bases.

Acids
1) taste sour
2) feel oily
3) litmus turns red
4) react w/ metals
5) conduct electricity
6. pH < 7
7. $[H^+] > [OH^-]$

Bases
1) taste bitter
2) feel slippery
3) litmus turns blue
4) conduct electricity
5) pH > 7
6. $[H^+] < [OH^-]$

2. Strong acids & bases dissociate (ionize) completely. Weak acids & bases only slightly dissociate (ionize).

3. On the periodic table, where are you most likely to find a strong acid? List the strong acids.

Groups 16 + 17 HCl, HI, HBr, HNO_3, H_2SO_4, $HClO_4$

4. On the periodic table, where are you most likely to find a strong base? List the strong bases.

Groups 1 + 2 $NaOH$, $LiOH$, KOH, $RbOH$, $CsOH$, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

5. Describe the differences between an Arrhenius and a Bronsted-Lowry acid and base.

Arrhenius acid contains H
Arrhenius base contains OH
Bronsted-Lowry acid is a H^+ donor
Bronsted-Lowry base is a H^+ acceptor

6. Identify the Bronsted-Lowry acid-base pairs in each of the following reactions. Label each substance.

a. NH_3 + H_2O <----> NH_4^+ + OH^-
BB BA Ca Cb

b. $HC_2H_3O_2$ + H_2O <----> $C_2H_3O_2^-$ + H_3O^+
BA BB Cb Ca

7. What are the formulas for hydroxide OH^- and hydronium H_3O^+?

8. If the hydronium concentration of a solution is 2.34×10^{-3} M, what is the pH?

$pH = -\log[H_3O^+]$ $pH = -\log 2.34 \times 10^{-3}$ $\boxed{pH = 2.63}$

9. If the concentration of HNO_3 is .00025M calculate the pH and pOH.

$pH = -\log[H^+]$ $pH = -\log .00025$ $\boxed{pH = 3.60}$

$pH + pOH = 14$ $pOH = 14 - pH$ $pOH = 14 - 3.60$ $\boxed{pOH = 10.4}$

10. What is the $[H^+]$ concentration of a solution with a pH of 2.687. What is the $[OH^-]$?

$[H^+] = 10^{-pH}$ $[H^+] = 10^{-2.687}$ $\boxed{[H^+] = .00206 \text{ M}}$

$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]}$ $[OH^-] = \frac{1 \times 10^{-14}}{.00206}$ $\boxed{[OH^-] = 4.85 \times 10^{-12} \text{ M}}$

11. Calculate the pH and the pOH for a 6.57×10^{-9} M solution of LiOH.

$pOH = -\log[OH^-]$ $pOH = -\log 6.57 \times 10^{-9}$ $\boxed{pOH = 8.18}$

$pH + pOH = 14$ $pH = 14 - pOH$ $pH = 14 - 8.18$ $\boxed{pH = 5.82}$

12. An acid + a base yields a Salt + H_2O. This type of reaction is called neutralization.
13. The process used to find the concentration of an acid or a base is a titration.

14. If 25 mL of 0.20 KOH were used to titrate 15 mL of H_2SO_4, what is the molarity of the acid? You must first complete and balance the equation. Show your work for the calculation. 2 KOH + ______ $H_2SO_4 \rightarrow 2H_2O + K_2SO_4$

25mL, .20M | 15mL, ?M

$$\frac{25\text{mL KOH}}{} \Big| \frac{1\text{L KOH}}{1000\text{mL KOH}} \Big| \frac{.2\text{mol KOH}}{1\text{L KOH}} \Big| \frac{1\text{mol } H_2SO_4}{2\text{mol KOH}} = \frac{.0025\text{mol } H_2SO_4}{.015\text{ L } H_2SO_4} = \boxed{.167\text{ M } H_2SO_4}$$

a. Which substance is the titration standard? KOH
b. What must be added to signal the end of the titration? indicator
c. How do we choose an appropriate substance to signal the end of a titration? depends on strong/weak acids/bases
d. What do we call it when the moles of acid = moles of base? equivalence point

15. Circle the strong base and put a box around the strong acid.
- $\boxed{\text{HCl}}$ NH3 CH3COOH NaOH

16. For the following types of titrations, give the pH range for the equivalence point:
- Weak acid-strong base pH > 7
- Strong acid-strong base pH around 7
- Strong acid- weak base pH < 7

17. 75.0 ml of .250M nitric acid, HNO_3, reacts with 25.5 ml of potassium hydroxide, KOH. What is the molarity of the base? Be sure to write a balanced chemical equation first. $HNO_3 + KOH \rightarrow H_2O + KNO_3$

75mL, .250M | 25.5mL, ?M

$$\frac{75\text{mL } HNO_3}{} \Big| \frac{1\text{L } HNO_3}{1000\text{mL } HNO_3} \Big| \frac{.250\text{mol } HNO_3}{1\text{L } HNO_3} \Big| \frac{1\text{mol KOH}}{1\text{mol } HNO_3} = \frac{.0188\text{ mol KOH}}{.0255\text{ L KOH}} = \boxed{.737\text{ M KOH}}$$

18. What is the molarity of $Ca(OH)_2$ solution if 30.5 ml of the solution is neutralized by 36.6 ml of .250 M HBr?

$$Ca(OH)_2 + 2HBr \rightarrow 2H_2O + CaBr_2$$

?M, 30.5mL | .250M, 36.6mL

$$\frac{36.6\text{mL HBr}}{} \Big| \frac{1\text{L HBr}}{1000\text{mL HBr}} \Big| \frac{.250\text{mol HBr}}{1\text{L HBr}} \Big| \frac{1\text{mol } Ca(OH)_2}{2\text{mol HBr}} = \frac{.00458\text{mol } Ca(OH)_2}{.0305\text{ L } Ca(OH)_2} = \boxed{.150\text{M } Ca(OH)_2}$$

B. Unit 9 -Solutions

Vocabulary: Review the following vocabulary. Look up and write the definition for any words you do not know.

Alloy	Insoluble	Solvation
Boiling point elevation	Miscible	Solvent
Colligative property	Molarity	Solute
Concentration	Saturated solution	Supersaturated solution
Dilution	Solubility	Suspension
Freezing Point depression	Soluble	Unsaturated solution
Immiscible	Solution	

Unit Objectives
- Discuss the factors affecting solubility and rate of dissolving one substance in another
- Use polarity of molecules to relate solubility of substances
- Describe the characteristics of a solution
- Calculate concentrations of solutions

19. Describe solute and solvent.
Solute gets dissolved
Solvent does the dissolving

20. List the factors that affect solubility.
1) agitation
2) particle size
3) temperature

21. In general, the solubility of most solid substances increases as temperature increases. The solubility of gases, however, decreases as temperature increases.

22. Describe the rule "Likes dissolves Like".

Same type of substance dissolves similar substances

Think of intermolecular forces; polar dissolves polar + ionic; nonpolar dissolves nonpola

23. What type(s) of compounds are soluble in water. ionic and polar

24. Describe the three types of solutions.

Saturated: Contains the maximum amount of solute

Unsaturated: Contains less than the maximum amount of solute

Supersaturated: Contains more than the maximum amount of solute

25. When you add more solvent to a solution, the solution becomes more dilute.

26. What unit do chemists use most often to describe concentration? Molarity (M) $\frac{mol\ solute}{L\ solution}$

27. Calculate the molarity of 0.205 L of solution that contains 63.8 g of NaOH.

$M = \frac{mol\ Solute}{L\ Solution}$ $\frac{63.8g\ NaOH \mid 1mol\ NaOH}{\mid 39.99711g\ NaOH} = \frac{1.60mol\ NaOH}{.205L} = \boxed{7.80\ M\ NaOH}$

28. **How would you prepare** 500 mL of 1.5 M NaCl from solid NaCl? Show any calculations needed.

$1.5M = \frac{mol\ Solute}{.500L}$ mol solute $= .75mol\ NaCl \mid \frac{58.44277g\ NaCl}{1mol\ NaCl} = 43.8g\ NaCl$

1. Calculate + measure 43.8g NaCl in a 500mL volumetric flask.
2. Add some water to dissolve the NaCl.
3. Add remaining water to the calibration mark on the flask.
4. Cap + invert to mix.

29. A .600 L sample of a 2.50 M solution of KI contains what mass of KI?

$2.50M = \frac{X}{.600L}$ $X = 1.5mol\ KI \mid \frac{166.0033g\ KI}{1mol\ KI} = \boxed{249g\ KI}$

30. What is the volume of 0.1250 M solution of $AgNO_3$ that contains 1.75 moles of solute.

$.1250M = \frac{1.75mol}{X}$ $\boxed{X = 14L}$

31. How many mL of 2.0 M KOH stock solution do you need to prepare 100 mL of 0.40 M KOH.

$M_1V_1 = M_2V_2$ $2.0M \times V_1 = .40M \times 100mL$

$\boxed{V_1 = 20mL}$

32. What is a colligative property?

a physical property of a solution that depends on the number of the dissolved solute particles (not identity)

33. What do colligative properties depend on?

# of dissolved solute particles

34. How is the boiling point of water affected when a solute is added? it increases How about the freezing point? it decreases

35. List in order which compound with equal concentrations has the greatest affect on raising the boiling point of a solution: NaCl, sugar ($C_{12}H_{22}O_{11}$), $CaCl_2$. **Why?** 1. $CaCl_2$ has the greatest affect b/c it is ionic + dissociates into 3 ions ($i=$
2. NaCl is next because it dissociates into 2 ions ($i=2$)
3. $C_{12}H_{22}O_{11}$ - covalent so it doesn't dissociate ($i=1$)

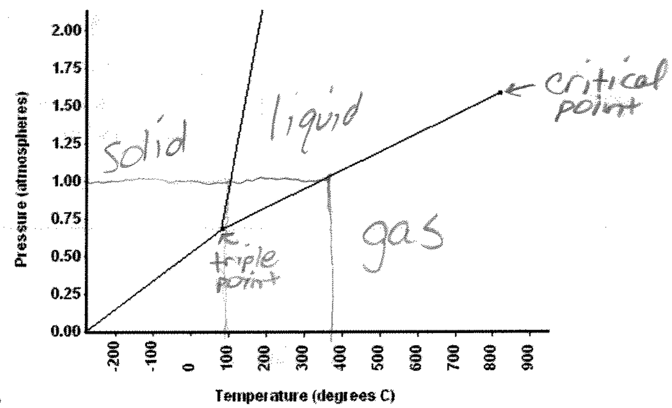
36. Explain why a solution has a lower freezing point than a pure solvent. Why does its boiling point also increase?
A solution has more particles than a pure solvent (i is more)
when a solute is added to a solvent the bp ↑ and the fp ↓

37. What is an electrolyte? substance that dissociates in water to form ions that will conduct electricity

38. What is a nonelectrolyte? substance that doesn't dissociate into ions so it won't conduct electricity

39. Give an example of each: Electrolyte- NaCl (ionic) Nonelectrolyte- CH_4 (covalent compounds)

40. Use the phase diagram to answer the following 4 questions:
   a. Label the following on the phase diagram below: Solid phase, liquid phase, gas phase, triple point, critical point.



   b. What is the normal (1 atm) melting point of this substance?
   ~100°C
   c. What is the normal boiling point of this substance?
   ~375°C
   d. What is the normal freezing point of this substance?
   ~100°C

C. Unit 10 Equilibrium

Vocabulary: Review the following vocabulary. Look up and write the definition for any words you do not know.

Dissociation equations	chemical equilibrium	Le Chatelier's Principle
reversible reaction	homogeneous equilibrium	K_{sp}
completion reaction	heterogeneous equilibrium	K_{eq}

Unit Objectives

- Describe the characteristics of chemical equilibrium
- Use LeChatelier's principle to predict the direction of reversible reactions
- Calculate K_{sp} and K_{eq}

41. Using the experimental data provided, determine the order of reaction with respect to each reactant, the rate law equation, and the overall order of reaction.
$CO(g) + Cl_2(g) \rightarrow COCl_2(g)$

Experiment	Initial Concentration (mol/L)		Initial Rate
	CO	Cl	(mol/L•s)
1	0.12	0.20	0.121
2	0.24	0.20	0.241
3	0.12	0.40	0.483

[CO] trials 1&2
$\left(\frac{.24}{.12}\right)^x = \frac{.241}{.121}$
$2^x = 2$
$x = 1$ 1st order

$[Cl_2]$ trials 1&3
$\left(\frac{.40}{.20}\right)^x = \frac{.483}{.121}$
$2^x = 4$
$x = 2$ so $[Cl_2]$ is 2nd order

$\boxed{\text{rate} = k[CO][Cl_2]^2}$
Overall order is 3 (1+2)

42. A double arrow signifies a reversible reaction, while a single arrow signifies a Complete reaction.

43. What causes a reaction to go to completion? The evolution of a gas or the formation of a precipitate.

44. Le Châtelier's Principle explains how an equilibrium system will respond to stress.

45. Describe chemical equilibrium. Give an example.
A state in which the forward & reverse reactions take place at equal rates
$PCl_{3(g)} + Cl_{2(g)} \leftrightarrow PCl_{5(g)}$

46. Write the equilibrium constant expression for $4HCl_{(g)} + O_{2(g)} \leftrightarrow 2Cl_{2(g)} + 2H_2O_{(g)}$
$K_{eq} = \frac{[products]}{[reactants]}$ $K_{eq} = \frac{[Cl_2]^2[H_2O]^2}{[HCl]^4[O_2]}$

47. If you calculate a small number (less than 1) for the constant expression above, what does that tell you?
the reactants are favored (the reverse rxn is favored)

48. At 773K, the reaction $2NO(g) + O_2(g) \leftrightarrow 2NO_2(g)$ produces the following concentrations: $[NO]=3.49 \times 10^{-4}$ M; $[O_2]=0.80M$; $[NO_2]=0.250M$. Write the equilibrium constant expression for the reaction, & calculate the value of the equilibrium constant.
$K_{eq} = \frac{[NO_2]^2}{[NO]^2[O_2]}$ $K_{eq} = \frac{(.250)^2}{[(3.49 \times 10^{-4})^2][.80]}$ $\boxed{K_{eq} = 6.41 \times 10^5 \text{ or } 641{,}000}$

49. For the reaction given, complete the following table: **C (s) + H_2O (l) + heat $\leftrightarrow$ CO (g) + H_2 (g)**

Stress applied	Shift left, shift right, or no change?	What happens to the concentration of CO?
Cooling	shift left	decreases
Adding water	shift right	increases
Adding a catalyst	no change	no change
Removing H_2	shift right	increases
Decreasing volume	no change	no change

50. For the reaction; **Heat + $H_{2(g)}$ + $I_{2(g)}$ $\leftrightarrow$ $2HI_{(g)}$**
A. How will an increase in temperature change the concentration of Hydrogen gas? it will decrease
B. How will an increase in pressure affect the system? no change (equal # of moles)
C. Which direction will the addition of Iodine gas shift the system? right What does this do to the concentration of Hydrogen gas? decreases

51. For the reaction **$N_2O_4(g)$ + heat $\leftrightarrow$ 2 $NO_2(g)$**
a. List 2 stresses that you could apply to the equilibrium system to increase the 2 NO_2 (g):
increase $[N_2O_4]$ decrease pressure / increase volume
add heat decrease $[NO_2]$
b. List 2 stresses that you could apply to the equilibrium system to increase the N_2O_4 (g):
decrease $[N_2O_4]$ increase pressure / decrease volume
remove heat increase $[NO_2]$

52. What is dissociation? Write and balance the equation for the dissociation of Na_3PO_4.
when a substance breaks down into its ions
$Na_3PO_4 \rightleftharpoons 3Na^+ + PO_4^{3-}$

53. Write the K_{sp} expression for the dissociation of AgBr (s) and calculate the concentration of [Ag] if the K_{sp} value for AgBr is 5.4×10^{-13}. What does this K_{sp} value tell you about the reaction?
$AgBr \rightleftharpoons Ag^+ + Br^-$
$K_{sp} = [Ag^+][Br^-]$
$x \quad x$
$5.4 \times 10^{-13} = x^2$
$x = \sqrt{5.4 \times 10^{-13}}$ $x = 7.35 \times 10^{-7}$ M, s
$[Ag^+] = 7.35 \times 10^{-7}$M
Small # so sparingly soluble

54. What is the K_{sp} for AgCl if the concentration of silver ions is 1.25×10^{-16} M?
$AgCl \rightleftharpoons Ag^+ + Cl^-$
$K_{sp} = [Ag^+][Cl^-]$
$K_{sp} = x^2$
$K_{sp} = (1.25 \times 10^{-16})^2$
$\boxed{K_{sp} = 1.56 \times 10^{-32}}$

D. Unit 10– Reaction Rates

Vocabulary: Review the following vocabulary. Look up and write the definition for any words you do not know.

Activated complex
Activation energy
Collision theory
Reaction rate
Transition state
Catalyst

Unit Objectives

- Distinguish between exo- and endothermic reactions and determine heat of reactions
- Identify and describe factors that influence the rate of reaction

55. List the factors that affect the RATE of a chemical reaction ***and*** tell HOW they affect the rate.

5 Factors that affect the reaction rate:	How the factors alter the rate:
nature of the reactants	Some elements are more reactive than others like group 1+2 metals
concentration	increase concentration increases collisions which increases rates
surface area	increase surface area by crushing increases the collisions which increases rates
temperature	increase temp. increases particle speed which increases collision + the rate
catalysts	increase rate of rxn by lowering the activation energy

56. What is a catalyst? How is an enzyme like a catalyst? How do catalysts work?
· A catalyst is a substance that increases the rate of a rxn w/o itself being consumed
· Enzymes are examples of catalysts
· Catalysts lower the activation energy

57. In order for a reaction to occur, the reactants must collide with enough energy and the correct orientation. This will create an activated complex which can form product.

58. The amount of energy needed for an effective collision is called the activation energy.

59. What happens to the rate of a chemical reaction over time? it slows down because you use up run out of reactants

60. In a chemical reaction that produces hydrogen 14.3 ml of gas was collected over a 20.0 second period. Calculate the rate of the reaction in ml/sec. rate = $\frac{14.3 \text{ mL}}{20.0 \text{ sec}}$ = $\boxed{.715 \text{ mL/sec}}$

E. Unit 8 – Energy and Chemical Changes

Vocabulary: Review the following vocabulary. Look up and write the definition for any words you do not know.

calorie
calorimeter
chemical potential energy
energy
enthalpy
enthalpy (heat) of combustion
enthalply (heat) of reaction
entropy
free energy
heat
Joule
Law of conservation of energy
Law of disorder
Molar enthalpy (heat) of fusion
Molar enthalpy (heat) of vaporization
Specific heat
Spontaneous process
Standard enthalpy (heat) of formation
Surroundings
System
Thermochemical equation
Thermochemistry
Universe

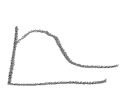
Unit Objectives

- Explain how changes in enthalpy, entropy, and free energy affect the spontaneity of chemical reactions and other processes.
- Write thermal equations and use them to calculate changes in enthalpy.
- Distinguish between exothermic and endothermic reactions
- Measure and calculate the energy involved in chemical changes.

61. Reactions that tend to be spontaneous have (negative, positive).
a. ΔH negative
b. ΔS positive
c. ΔG negative

62. Define entropy, enthalpy and free energy.
entropy – measure of disorder (S)
enthalpy – measure of heat content at constant pressure (H)
free energy – energy that is available to do work (G)

63. Describe an endothermic and exothermic reaction.

endothermic – heat as a reactant; +ΔH; product energy is greater than reactant energy

exothermic – heat as a product; -ΔH; product energy is less than reactant energy

64. In nature, do things tend to become more organized or more disordered? How is this related to entropy?

more disordered and release energy

Entropy (ΔS) is the measure of disorder

65. The enthalpy of the products is 255 kJ and the enthalpy of the reactants is 335 kJ. Calculate the change in enthalpy and determine if the reaction is exothermic or endothermic.

ΔH_{rxn} = products − reactants

ΔH_{rxn} = 255 kJ − 335 kJ = $\boxed{-80\text{ kJ}}$

exothermic b/c the ΔH is negative

66. Predict the sign of ΔS_{system} for the following changes and explain your answer:

a. $ClF\ (g) + F_2\ (g) \rightarrow ClF_3\ (g)$ − ΔS b/c it is less disordered on the product side

b. $C_{10}H_8\ (l) \rightarrow C_{10}H_8\ (s)$ −ΔS b/c it is less disordered on the product side

67. Given ΔH_{system}, T, and ΔS_{system}, determine if the following process is spontaneous or non-spontaneous: ΔH_{system} = -75.9 kJ, T = 273 K, and ΔS_{system} = 138 J/K. $\Delta G = \Delta H - T\Delta S$

ΔG = − 75.9 kJ − 273 K(.138 kJ/K)

ΔG = −114 kJ
Spontaneous b/c ΔG is negative

68. Is the following reaction spontaneous at 456 K? If not, is it spontaneous at some other temperature? Explain your answer.

$N_2\ (g) + 2\ O_2\ (g) \rightarrow 2\ NO_2\ (g)$ ΔH = 68 kJ and ΔS = -122 J/K $\Delta G = \Delta H - T\Delta S$

ΔG = 68 kJ − 456 K(−.122 kJ/K)

ΔG = 124 kJ

nonspontaneous b/c ΔG is positive

No, it will never be spontaneous when the ΔH is positive + the ΔS is negative.

F. Unit 7-Gas Laws

Vocabulary: Review the following vocabulary. Look up and write the definition for any words you do not know.

Avogadro's Principle	Ideal gas constant	Pascal
STP	Barometer	Pressure
Combined gas law	Dalton's law	Dipole-dipole forces
Molar volume	Diffusion	Dispersion forces
Ideal gas law	Kinetic-molecular theory	Hydrogen bond

Unit Objectives

- Explain the concept of an ideal gas and perform calculations
- Use stoichiometry to convert between substances in chemical reactions
- Use the kinetic-molecular theory to describe the behavior of gases

69. What are the four variables that describe a gaseous system?

1. Pressure 2. Volume 3. Temperature 4. n (moles)

70. Temperature must *always* be in Kelvin when calculating gas law problems.

a. 24°C = 297 Kelvin

b. 392 K = 119 Celsius

71. Standard pressure = 1 atm = 101.3 kPa = 760 mmHg

72. Standard temperature = 273 K = 0 degrees Celsius.

73. When the amount of gas in a container increases the pressure increases **because**.... there are more particles so more collisions

74. When the temperature of a gas increases its volume will increase if the pressure is kept constant **because**....

they are directly related

75. When you increase the volume on a sample of gas the pressure will decrease **because**...

they are indirectly related more volume means less collisions

76. Answer the following questions with INVERSELY or DIRECTLY.

a. How are pressure and temperature related? directly

b. Pressure and Volume? indirectly

c. Volume and Temperature? directly

77. Write the formula for the combined gas law. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

78. 150 mL of oxygen has a pressure of 752 mm Hg at 22°C. Calculate its volume at STP.

$\frac{752\,mm\,Hg \times 150\,mL}{295\,K} = \frac{760\,mmHg \times V_2}{273\,K}$ $\boxed{V_2 = 137\,mL}$

79. If 51.30 Liters of a gas is collected at a pressure of 59.0 kPa, +290K what volume will the same gas occupy at 101.3 kPa if the temperature ~~remains constant~~? drops to 274K

$\frac{59.0\,kPa \times 51.30\,L}{290\,K} = \frac{101.3\,kPa \times V_2}{274\,K}$ $\boxed{V_2 = 28.2\,L}$

80. If the volume of a gas is 26ml at 24.8 degrees Celsius +1.3atm what ~~volume~~ temp. in Celsius would the balloon occupy at ~~88.8 degrees Celsius~~? 1.5atm 15mL?

$\frac{1.3\,atm \times 26\,mL}{297.8\,K} = \frac{1.5\,atm \times 15\,mL}{T_2}$ $T_2 = 198\,K - 273 = \boxed{-75°C}$

81. How many moles of a gas will occupy 2.50L at STP?

$\frac{2.50\,L \mid 1\,mol}{\mid 22.4\,L} = \boxed{.112\,mol}$ or $1\,atm \times 2.50\,L = n \times .0821 \times 273\,K$

$n = .112\,mol$

82. Calculate the volume that 3.60 grams of H_2 gas will occupy at STP.

$\frac{3.60g\,H_2 \mid 1\,mol\,H_2}{\mid 2.01588g\,H_2} = \frac{1.79\,mol\,H_2 \mid 22.4\,L}{\mid 1\,mol\,H_2} = \boxed{40.1\,L}$

83. If a 0.500 g sample of gas occupies 255 mL at 25°C and a pressure of 1.10 atm, what is the molar mass of the gas?

$PV = nRT$

$1.1 \times .255\,L = n \times .0821 \times 298\,K$ OR $M = \frac{mRT}{PV}$ $M = \frac{.500g \times .0821 \times 298\,K}{1.10 \times .255\,L}$ $\boxed{M = 43.6\,g/mol}$

$n = .0115\,mol$ $\frac{.5g}{.0115\,mol} = \boxed{43.5\,g/mol}$

84. Use the reaction shown to calculate the mass of iron that must be used to obtain .500L of hydrogen at 24.3 degrees Celsius and 100.0 kPa of pressure. **3Fe + 4H₂O → Fe₃O₄ + 4H₂** $PV = nRT$

$PV = nRT$ $.987\,atm \times .500\,L\,H_2 = n \times .0821 \times 297.3\,K$

$\frac{100.0\,kPa \mid 1\,atm}{\mid 101.3\,kPa} = .987\,atm$ $n = \frac{.0202\,mol\,H_2 \mid 3\,mol\,Fe \mid 55.847g\,Fe}{\mid 4\,mol\,H_2 \mid 1\,mol\,Fe} = \boxed{.846g\,Fe}$

85. What is an intermolecular force? How do they affect the melting point of different substances?

Forces of attraction or repulsion which act between neighboring particles.

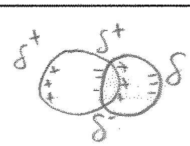
The stronger the intermolecular force, the higher the boiling point

LDF < dipole-dipole < hydrogen bond < ionic bond

86. What kinds of molecules exhibit hydrogen bonding and how does it contribute to the relatively high boiling point for water?

Molecules where H is bonded to a very electronegative element like N, O, F

The water is very polar which holds the molecules in a liquid state until a high temp. is reached

87.	London Dispersion Forces	Dipole-Dipole	Hydrogen Bonds
Definition 	weak forces that result from temporary shifts in the density of electrons in the electron cloud	attractions between oppositely charged regions of polar molecules δ⁻ δ⁺ δ⁻ δ⁺	Hydrogen bonded to a small highly electronegative atom w/at least 1 lone e⁻ pair (N, O, F)
This type of force would be found between what type of molecules?	nonpolar molecules	polar molecules	hydrogen bonded to a very electronegative element (N, O, F)
Rank these three forces from strongest to weakest and **explain why.**	weakest (temporary dipole)	permanent dipoles	very strong

88. Circle the chemicals that are soluble in water. Cross out the ones that are not.

~~Fe (iron)~~ MgCl$_2$ ~~C_5H_{10}~~ ~~SiO_2~~ O=Si=O

element ionic nonpolar symmetrical nonpolar symmetrical

G. Multiple Choice Review Questions: Choose the best answer to complete each question.

1. Which of the following has the lowest freezing point?
   a. KBr ionic i=2
   b. CCl_4 nonpolar i=1
   c. H_2O hydrogen bond polar
   d. NCl_3 dipole dipole polar
2. Which of the following reactions has a decrease in entropy? $-\Delta S$
   a. $H_2O\ (l) \rightarrow H_2O\ (g)$
   b. $2\ O_3\ (g) \rightarrow 3\ O_2\ (g)$
   c. $CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$
   d. $3\ H_2\ (g) + N_2\ (g) \rightarrow 2\ NH_3\ (g)$ less disorder
3. Which of the following has the highest boiling point?
   a. Ammonia (NH_3) polar w/hydrogen bond
   b. Water (H_2O) polar w/hydrogen bond
   c. Lithium fluoride (LiF) ionic i=2
   d. Methane (CH_4) nonpolar
4. Solid sodium hydroxide is added to water in a sealed container. Which of the following statements is true?
   a. Entropy and the total energy remain constant. energy stays the same
   b. Entropy increases, total energy is constant. NaOH(s) + H$_2$O is more disordered
   c. Entropy decreases, total energy is constant.
5. For this reaction, which of the following statements is false? $H_3O^+ + CO_3^{2-} \leftrightarrow HCO_3^- + H_2O$ (BA BB CA CB)
   a. The carbonate ion is a Bronsted base
   b. The bicarbonate ion is a conjugate acid
   c. The hydronium ion is a Bronsted acid
   d. The water is the conjugate acid F it is the cb
6. Which mixture is used to prepare 500 mL of a 0.20 M solution of sodium sulfate? Na_2SO_4
   a. 14.2 g of solute dissolved to make 500 mL of solution $M = \frac{mol}{L}$ total solution volume is 500mL
   b. 14.2 g of solute dissolved in 500 mL of water
   c. 28.4 g of solute dissolved in 1 L of water
   d. 14.2 g of solute dissolved in 500 g of water

   $.2M = \frac{x}{.500L}$ $x = .10 mol\ Na_2SO_4 \times \frac{142.03714 g\ Na_2SO_4}{1 mol\ Na_2SO_4} = 14.2 g\ Na_2SO_4$
7. Which of the following has the lowest pH?
   a. 0.10 M HCl strong acid means lower pH
   b. 0.10 M CH_3COOH weak acid
   c. 0.10 M H_2CO_3 weak acid
   d. They are all the same
8. A solution of a monoprotic strong acid has a pH of 2.10. What is the concentration of the acid?
   a. 0.00794 $[H^+] = 10^{-2.10}$
   b. 0.00931
   c. 0.110
   d. 0.0202
9. For a strong acid-weak base titration, which indicator would be most appropriate? SA wb 0 6.5 7 14
   a. Crystal violet (color change pH 0.5-1.5)
   b. Methyl red (color change pH 5.0-5.7)
   c. Bromthymol blue (color change pH 6.0-7.3)
   d. Alizarin yellow (color change pH 10.3-11.8)
10. Which of the following does NOT contain hydrogen bonds? H bonded to N, O, F
    a. Water, H_2O
    b. Ammonia, NH_3
    c. Acetic acid, CH_3COOH
    d. Dimethyl either, CH_3OCH_3 H bonded to C
11. If the volume of a sample of gas in a piston is decreased to one-third of its original value at constant temperature, which of the following will increase proportionally? ↓1/3
    a. Celsius temperature
    b. Pressure P+V indirectly related
    c. Velocity of the molecules
    d. kinetic energy
12. A gas sample in a piston container has a volume of 2.0 liters at 1.0 atm and 27 °C. The temperature is changed such that the volume is decreased to 1.2 liters and the pressure is increased to 5.0 atm. What Kelvin temperature is needed to produce this change? $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ $\frac{1atm \times 2.0L}{300K} = \frac{5atm \times 1.2L}{T_2}$
    a. 15.0K
    b. 273 K
    c. 623 K
    d. 900K
13. A 1.0 liter flask is filled with a mixture of two gases at 20 °C until a pressure of 14.43 atm is established. If 0.40 grams of the mixture is hydrogen, how many moles are there of the other gas? PV=nRT
    a. 0.20 moles
    b. 0.30 moles
    c. 0.40 moles
    d. 0.50 moles

    $14.43 atm \times 1.0L = n \times .0821 \times 293K$ $n = .5999 mol$

    $.5999 mol = n + .198 mol\ H_2$ $n = .4019$ mol other gas

    $\frac{.40g\ H_2}{} \left| \frac{1 mol\ H_2}{2.01528g\ H_2} \right. = .198 mol$
14. A 1.5 g sample of a gaseous hydrocarbon has a volume of 820 mL when measured at 227 °C and 2.50 atm. Which of the following is the formula for the gas? PV=nRT
    a. CH_4 12.0111+4(1.00794)=16.0g/mol $2.50 atm \times .820L = n \times .0821 \times 500K$
    b. C_2H_6 2(12.0111)+6(1.00794)=30.1 g/mol $n = .0499 mol$
    c. C_3H_8 3(12.0111)+8(1.00794)=44.097g/mol
    d. C_4H_{10} 4(12.0111)+10(1.00794)=58.1g/mol

    $\frac{1.5g}{.0499 mol} = \boxed{30.1 g/mol}$

15. Which of the following water solutions has the lowest freezing point?

a. 0.3 m sucrose polar so i=1
b. 0.20 m $CaCl_2$ ionic w/ 3 ions so i=3 greater effect
$CaCl_2 \rightarrow Ca^{2+} + 2Cl^-$
c. 0.20 m NaCl ionic w/ 2 ions
d. 0.20 m NH_4Cl ionic w/ 2 ions

16. Which applies to the colligative properties of solutions?

I. They depend on the specific kind of particles in the solute.
II. They affect the boiling point of a solution.
III. They affect the freezing point of a solution.

a. II only
b. III only
c. II and III only
d. I, II, and III

17. Which of the following will increase the molar solubility of an ionic salt in water? ability to dissolve

a. Stir the solution
b. Add more solute
c. Crush the solute
d. Heat the solution

18. Which of the following affects the boiling point of a liquid?

a. The intermolecular forces need to know for the i
b. The volume
c. The mass
d. The size of the particles

19. For the exothermic reaction $C_3H_8 + 5O_2 \rightarrow 3\ CO_2 + 4\ H_2O$, which of the following is true at all temperatures?

I. $\Delta G < 0$ combustion rxn so it's spontaneous; ΔG would be −
II. $\Delta S > 0$ more disorder so +ΔS
III. $\Delta H > 0$ combustion rxn, so ΔH should be − & this isn't

a. I only
b. II only
c. III only
d. I and II only

20. In the following reaction ΔH_f is zero for _______. $Ni\ (s) + 2\ CO\ (g) + 2\ PF_3\ (g) \rightarrow Ni(CO)_2(PF_3)_2\ (l)$

a. Ni (s) elements + diatomics are always zero
b. CO (g)
c. PF_3 (g)
d. Both CO (g) and PF_3 (g)

21. The value of ΔH for the reaction below is -72 kJ. _____ kJ of heat are released when 80.9 grams of HBr is formed.

$H_2\ (g) + Br_2\ (g) \rightarrow 2\ HBr\ (g)$

$\frac{80.9\text{g HBr}}{} \cdot \frac{1\text{mol HBr}}{80.9119\text{g HBr}} \cdot \frac{72\text{kJ}}{2\text{mol HBr}} = 35.99\text{kJ}$

a. 144
b. 72
c. 36
d. -72

22. The value of ΔH for the following reaction is -3351 kJ: $2\ Al\ (s) + 3\ O_2\ (g) \rightarrow 2\ Al_2O_3\ (s)$. The value of ΔH_f for Al_2O_3 (s) is _______ kJ. $\Delta H_{rxn} = \Sigma \text{products} - \Sigma \text{reactants}$

$-3351\text{ kJ} = 2(x) - 0$

$-1675.5 = x$

a. -3351
b. -1676
c. -32.86
d. +3351

23. The enthalpy of formation of a compound is -184 kJ/mol, and the products of its combustionhave a total enthalpy formation of -1356 kJ. What is the enthalpy of combustion of this compound?

$\Delta H_{rxn} = \Sigma \text{products} - \Sigma \text{reactants}$

$\Delta H_{rxn} = -1356 - -184$

a. -1172
b. -150
c. +1172
d. -1892

24. Which of the following should have the lowest boiling point? LDF < dipole-dipole < H-bond < ionic

a. PH_3 polar so dipole-dipole
b. H_2S polar so dipole-dipole
c. SiH_4 nonpolar so LDF
d. H_2O polar H-bond

25. Which of the following derivatives of ethane has the highest boiling point?

all the same compound, just w/ a different halogen, so since we want highest we need the one w/ the most mass

a. C_2Br_6
b. C_2F_6
c. C_2I_6
d. C_2Cl_6

26. Which of the following has dispersion forces as its only intermolecular force?

a. CH_4 nonpolar – LDF
b. HCl polar dipole-dipole
c. NaCl ionic
d. CH_3Cl polar dipole-dipole

27. The predominant intermolecular force in $CaBr_2$ is _______.

metal nonmetal so ionic bond

a. London- dispersion forces
b. Ion-dipole forces
c. Dipole-dipole forces
d. Ionic bonding

28. Of the following, ______ is an exothermic process.

a. Melting (s) + heat → (l) heat on product side
b. Subliming (s) + heat → (g)
c. Freezing (l) → (s) + heat
d. Boiling (l) + heat → (g)