109 Things to Know to Pass the Chemistry Midterm

(NOTE: this is NOT all inclusive- you must study old tests, review sheets, and other review materials)

1. ***Protons*** are positively charged (+) with a mass of 1 amu.
   Example: Which has the greatest nuclear charge? *Cl-35* *Ar-40* *K-39* *Ca-40*
   17 18 19 20

Neutrons have no charge and a mass of 1 amu.

3. ***Electrons*** are small and are negatively charged (-) with a mass of almost 0 amu..

4. Protons & neutrons are in an atom's nucleus (***nucleons***).
   Which has the greatest number of nucleons? Sn-119 *Sb-122* *Te-128* *I-127*

5. Electrons are found in "clouds" (***orbitals***) around an atom's nucleus.
   Where is most of the mass of an atom found? nucleus
   Where is most of the size (volume) of an atom found? orbitals

6. The ***mass number*** is equal to an atom's number of protons and neutrons added together.
   What is the mass number of an atom with 18 protons and 22 neutrons? 40 u

7. The ***atomic number*** is equal to the number of protons in the nucleus of an atom.
   Which has the greatest atomic number? *S* *Cl* *Ar* *K*

8. The ***number of neutrons*** = mass number – atomic number. *Which correctly represents an atom of neon containing 11 neutrons?* ^{11}Ne ^{21}Ne ^{20}Ne ^{22}Ne

9. In a neutral atom the number of protons = the number of electrons.

10. ***Isotopes*** are atoms with equal numbers of protons, but differ in their neutron numbers.
    Two isotopes of the same element will have the same number of
    neutrons and electrons, *neutrons and nucleons,* *protons and nucleons,* *protons and electrons*

11. Cations are *positive* (+) ions and form when a neutral atom *loses* electrons. They ***are*** *smaller* than their parent atom.
    Which of the following will form an ion with a smaller radius that that of its atom?
    Cl *N* *Br* *Ba* metal

12. ***Anions*** are negative ions and form when a neutral atom *gains* electrons. They are larger than their parent atom.
    Which electron configuration is correct for a fluoride ion?2–7 *2–8* *2–8–1* *2–6*

13. **Ernest *Rutherford's*** *gold foil experiment* showed that an atom is mostly empty space with a small, dense, positively charged nucleus.

J.J. Thompson discovered the electron and developed the "plum-pudding" model of the atom.

+ - + -
\+ - + - +
\- + - +

Positive & negative particles spread throughout entire atom.

15. ***Dalton's*** model of the atom was a solid sphere of matter that was uniform throughout.

16. The ***Bohr Model*** of the atom placed electrons in "planet-like" orbits around the nucleus of an atom.

17. The current, ***wave-mechanical model*** of the atom has electrons in "clouds" (orbitals) around the nucleus.

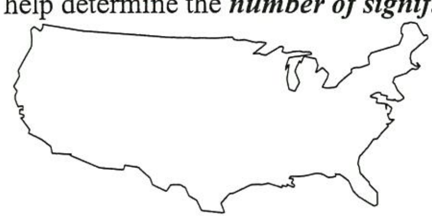
18. Electrons can be excited to jump to higher energy levels. They emit energy as light when they fall from higher energy levels back down to lower (***ground state***) energy levels. ***Bright line spectra*** are produced.

19. ***Elements*** are pure substances composed of atoms with the same atomic number. They cannot be decomposed.
    A compound differs from an element in that a compound
    Has a homogeneous composition *has one set of properties*
    Has a heterogeneous composition *can be decomposed*

20. ***Binary compounds*** are substances made up of only *two* kinds of atoms. "Ternary" compounds contain three (or more) kinds of atoms. *Which substance is a binary compound?*
    NH_3 *Ammonia* Mg *magnesium* *potassium nitrate* KNO_3 *methanol* CH_3OH

21. ***Diatomic molecules*** are elements that form two atom molecules in their natural form at STP.
    Which element is a diatomic liquid at STP? Chlorine *fluorine* *bromine* *iodine*

22. Use this diagram to help determine the ***number of significant figures*** in a measured value… OR NAS-D



Pacific **Atlantic**

If the decimal point is ***present***, start counting digits from the ***Pacific*** (left) side, starting with the first non-zero digit.
0.003100 (…4.. sig. figs.)
If the decimal point is absent, start counting digits from the Atlantic (right) side, starting with the first non-zero digit.
31,400 (……3.sig. figs.)

23. When multiplying or dividing measurements, final answer must have as many digits as the measurement with the fewest number of digits. When adding or subtracting, use place value.

What is the density of the object measured in lab by the displacement of water according to The data below:

Mass of object: 23.6 g
Volume of water: 15.0 mL
Volume of water + object: 18.2 mL

$\begin{array}{r} 18.2 \\ -15.0 \\ \hline 3.2 \end{array}$ $\frac{23.6}{3.2} = 7.4\ g/mL$

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

25. ***Heterogeneous mixtures*** have discernable components and ***are not*** uniform throughout.

Air is classified chemically as a(n)

Substance *compound* *element* *mixture*

26. Isotopes are written in a number of ways: C-14 is also Carbon-14, and is also

$^{14}_{6}C$

atomic number = ...6... *mass number* = ...14...

27. The average atomic mass is the weighted average mass of all the known isotopes of an element.

Find the average atomic mass of lithium if 7.4 % are ^{6}Li and 92.6% are ^{7}Li. $\frac{(7.4)(6)+(92.6)(7)}{100} = 6.926$

28. The distribution of electrons in an atom is its ***electron configuration***.

29. Electron configurations are written in the bottom center of an element's box on the periodic table in your reference tables. The outermost electrons are the valence electrons.

Al
13
2-8-3

2 = # of electrons in ...1st energy level
8 = # of electrons in ...2nd...
3= # of electrons in ...3rd.......

30. Use the **mole *map*** to help you solve conversions between moles, grams, numbers of molecules/atoms at STP.

Given the reaction $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$,

How many grams of CO2 are produced if 5.5 moles are produced? 5.5 mol = $\frac{x}{44}$

1 gram *44grams* *.125 grams* *242 grams*

31. ***Electron dot model*** is a way of representing the valence electron of an atom.

$\dot{X}$ *represents the electron-dot symbol of this element* C O B N

32. Energy is *absorbed* when a chemical bond breaks. Energy is *released* when a chemical bond forms. The greater the energy, the more stable the bond that forms. *Which of the following releases energy?*

$H_2 \rightarrow H + H$ $2NaCl \rightarrow 2Na + Cl_2$ $2Ca + O_2 \rightarrow 2CaO$ $2LiOH \rightarrow 2LiH + O_2$

33. Polyatomic ions (Table E) are groups of atoms, **covalently** bonded together, with an overall charge.

Nitrate: ...NO_3^{-1}..., NH_4^+:...ammonia... sulfite:...SO_3^{2-}..., etc.

Which of the following contains both ionic and covalent bonds?

mg2NM *NaOH* *CH_3OH* *NaCl* *Cl_2*

34. ***Coefficients*** are written in front of the formulas of reactants and products to balance chemical equations. They give the ratios of reactants and products in a balanced chemical equation.

...2...*Na* + ...*Cl_2* → ...2......*NaCl*

35. Chemical formulas are written so that the charges of cations and anions neutralize (cancel) one another.

calcium phosphate: Ca^{2+} PO_4^{3-} = ...$Ca_3(PO_4)_2$

36. When naming binary ionic compounds, write the name of the positive ion (cation) first, followed by the name of the negative ion (anion) with the name ending in "-ide." *$CaCl_2$* calcium chloride *MgS* Magnesium sulfide

37. When naming compounds containing polyatomic ions, keep the name of the polyatomic ion the same as it is written in Table E. *NH_4Cl* ammonium chloride *Dimercury (I) nitrate* Hg_2^{+2} NO_3^{-1} $Hg_2(NO_3)_2$

38. **Roman numerals** are used to show the positive oxidation number of the cation if it has more than one positive oxidation number

FeO: iron(II) oxide... *Nickel (III) sulfate*: ......Ni^{+3} $SO_4^{(-2)}$ $Ni_2(SO_4)_3$

39. ***Physical changes*** do not form new substances. 3

They merely change the appearance of the original material. (The melting of ice) $H_2O (s) \rightarrow H_2O (l)$

40. ***Chemical changes*** result in the formation of new substances. *Which process is an example of a chemical change?*

the melting of ice *the electrolysis of water* *the boiling of water*

s→l l→s

41. ***Reactants*** are on the left side of the reaction arrow and ***products*** are on the right.

42. ***Temperature*** is a measure of average kinetic.

Which sample has the highest average kinetic energy? 273K

H_2O (l) at 0°C H_2O (s) at 0°C CO_2 (g) at STP Mg (s) at 298K

43. ***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: $CH_4 (g) + 2\ O_2 (g) \rightarrow 2\ H_2O (g) + CO_2 (g)$ + heat — exo

What is the overall result when CH_4 (g) burns according to this reaction?

Energy is absorbed.

Energy is released?

44. *Only* coefficients can be changed when balancing chemical equations!

Given the unbalanced equation: $4Al + 3O_2 = 2Al_2O_3$

When this equation is balanced using the smallest whole numbers, what is the coefficient of Al?

1 2 3 4

45. ***Synthesis reactions*** occur when two or more reactants combine to form a single product.

$2 Na_{(s)} + Cl_{2(g)} \rightarrow$ ……2NaCl

46. ***Decomposition reactions*** occur when a single reactant forms two or more products

$CaCO_{3(s)} \rightarrow CaO_{(s)}$ + …O_2

47. ***Single replacement reactions*** occur when one element replaces another element in a compound.

Which equation below represents a reaction classified as a "single replacement" reaction?

$2\ H_2 + O_2 \rightarrow 2\ H_2O$

$Pb(NO_3)_2 + K_2CrO_4 \rightarrow 2\ KNO_3 + PbCrO_4$

~~$HCl + NaOH \rightarrow H_2O + NaCl$~~

$Cu + Zn(NO_3)_2 \rightarrow Zn + Cu(NO_3)_2$

$CaCO_3 \rightarrow CO_2 + CaO$

$K_2S + Pb(C_2H_3O_2) \rightarrow$

48. ***Double replacement reactions*** occur when two compounds react to form two new compounds.

Potassium sulfide is mixed with lead acetate. Which of the following products is expected?

$PbSO_4$ K_2S K_3PO_4 PbS $Pb(C_2H_3O_2)_2$

49. The masses (and energy and charge) of the reactants in a chemical equation is always equal to the masses (and energy and charge) of the products. "***Law of Conservation of Mass(and Energy)***."

50. The gram formula mass (molar mass) of a substance is the sum of the atomic masses of all the atoms in it. = 130.2

H_2SO_4 = 132.2g/mole

2 x H = 2 x ..1.0...g = ..1.0..g 1 x S = 1 x 32.1..g = 32.1..g 4 x O = 4 x..16...g = 64.....g

51. Know how to calculate the percentage composition of a compound. (Formula is on Table T.)

Find the percent by mass of oxygen in $CaCO_3$.

52. ***6.02×10^{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)

53. Know how to convert a molecular formula into an empirical formula.

A compound has the molecular formula N_6O_{12}. Find its empirical formula.

N_3O_6 NO_2 N_2O_4 6 N_2O

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

55. 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) solid H_2O (l)

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

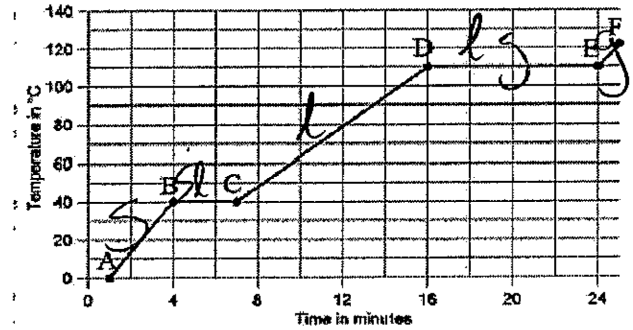
In what region of the graph below would you only find molecules with definite shape and volume? AB(noxA pase

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

58. ***Gases*** have widely-spaced particles that are in random motion (collide with container to create pressure).

59. ***Gases*** are easily compressed and have no definite shape or volume.
In what region of the graph below would you only find a sample with no definite shape or volume? EF

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



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

BC

61. Substances that ***sublime*** turn from a solid directly into a gas. They have very weak attractive forces. (examples include CO_2 & I_2)

62. As they evaporate, liquids become gases, which create vapor pressure. (Reference Table H). As temperature increases, vapor pressure increases. *This liquid on Reference Table H has the **weakest** attractive forces:*
Propanone *ethanol* *water* *acetic acid*

63. "***STP***" means "***S***tandard ***T***emperature and ***P***ressure." *Reference Table B*
These conditions define STP *P = ..1..atm* *T = 273 K*

64. Degrees Kelvin = C + 273
Room temperature = 25°C = 298 K *Boiling point of helium = 4 K = -269 °C*

65. Heat is a transfer of energy from a material at higher temperature to one at lower temperature.
When an ice pack is applied to a bruised arm, ... transfers from arm to ice pack

66. Use this formula to calculate heat absorbed/released by substances. $q = mc\Delta t$

q = heat absorbed or released (Joules)
m = mass of substance in grams
c = specific heat capacity of substance (J/gC) ... for water it's 4.18 J/g C.
Δt = temperature change in degrees Celsius

What is the total number of joules of heat energy absorbed by 12 grams of water when it is heated from 30°C to 40°C? $(12)(4.18)(10) = 501.6$

67. The heat absorbed or released when 1 gram of a substance changes between the solid and liquid phases is the substance's ***heat of fusion***. (*Reference Table B*: 334 J/g for water). *How many joules are required to melt 15 g H_2O (s)?* $Q = (15)(334)$ $= 5010$ J

68. The heat absorbed or released when 1 gram of a substance changes between the liquid and gaseous phases is the substance's ***heat of vaporization***. 2260J/g (*Reference Table B*)
How many joules are required to boil 120 g H_2O (l)? $(120)(2260) = 271200$ J

69. *Always use Kelvins* for temperature when using the ***combined gas law***.

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

$$\frac{50(1\ atm)}{273K} = \frac{(2.4atm)X}{240K}$$

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.

70. As the ***pressure*** exerted on a gas increases, the ***volume*** decreases proportionally. $(25)(1.2) = (.80)X$ $X = 7.5$ mL
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.

71. As the ***pressure*** on a gas increases, ***temperature*** increases. $\frac{220}{373} = \frac{x}{573}$ $X=220$ no change
A sample of gas exerts a pressure of 220. kPa at 373 K. Find the pressure at 373 K at constant volume.

72. As the ***temperature*** of a gas increases, ***volume*** increases. $\frac{15}{273} = \frac{x}{323}$ $X = 17.74$ mL
15 mL of oxygen gas is collected at 0°C. Find the volume at 50°C at constant pressure. Temp must be in Kelvin!

73. ***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*

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

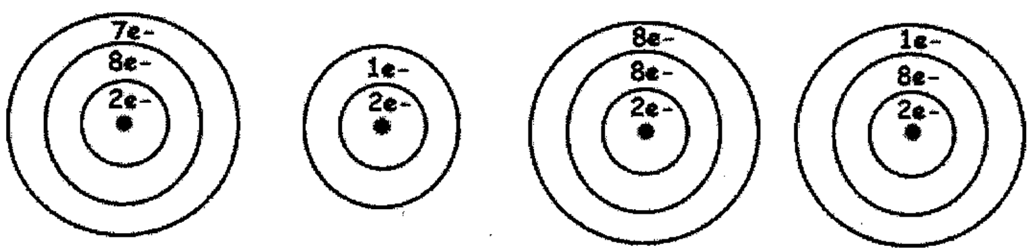
75. Mixtures may be separated by several physical means:
Distillation separates mixtures with different boiling points. *Fractional distillation is a common method to separate and collect:* *Hydrocarbons* *Ionic solids* *Metals* *Precipitates*
Filtration separates mixtures of solids and liquids.
*What would collect **in filter paper** if a mixture of NaCl (aq) and $CaCO_3$ (s) were poured through?*
Chromatography can also be used to separate mixtures of liquids and mixtures of gases.

76. ***The Periodic Law*** states that the properties of elements are periodic functions of their *atomic numbers*.
Elements are arranged on the modern periodic table in order of increasing ..atomic #
77. ***Periods*** are horizontal rows on the Periodic Table.
In which energy level are the valence electrons of the elements in Period 3 found? 3
78. ***Groups*** are vertical columns on the Periodic Table.
Which group on the periodic table contains a solid, liquid, and gas(es)? 17 (halogens)
79. ***Metals*** are found left of the "**staircase**" on the Periodic Table and at the bottom, ***nonmetals*** are above it and at the top, and ***metalloids*** border it.
Which of the following Group 14 elements has the greatest metallic character?
Carbon *silicon* *germanium* *tin* (furthest down a group)
80. Complete and memorize this chart.

Metals	Malleable and ductile	All solids except ……..	Lustrous	Good conductors of heat & electricity	……. ionization energy and electroneg.	Tend to form …… ions
Nonmetals	Brittle when solid	Mostly gases at STP	Dull	Good insulators	…….. ionization energy and electroneg.	Tend to form ….. ions

81. ***Noble gases*** (Group 18) are unreactive and stable due to the fact that their valence level of electrons is completely filled.
82. ***Ionization energy*** increases as you go up and to the right on the Periodic Table.
Which element among the diagrams below has the lowest ionization energy? D
83. ***Atomic radii*** *decrease* left to right across a period due to increasing nuclear charge.
Which period 3 element among the diagrams below has the largest radius? D
84. ***Atomic radii*** *increase* as you go down a group due to increased electron energy levels.
Which alkali metal among the diagrams below has the largest radius? D
85. ***Electronegativity*** is a measure of an element's attraction for electrons.
Which of the following atoms has the greatest tendency to attract electrons?
calcium *carbon* *copper* *chlorine*
86. Electronegativity *increases* as you go up and to the right on the Periodic Table.
Which element among the diagrams below has the greatest electronegativity? A
87. The elements in Group 1 are the ***alkali metals***; those in Group 2 are the ***alkaline earth metals***.
Which atom below represents the alkali metal of period 2? B
88. The elements in Group 17 are the ***halogens***.
Which element among the diagrams below is a halogen? A
89. The elements in Group 18 are the ***noble gases***.
Which element among the diagrams below is a noble gas? C

Element A **Element B** **Element C** **Element D**



317K=°C+273
90. Use ***Table S*** to compare and look up the properties of specific elements.
The freezing point of phosphorus is ..44..*°C*
91. 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*
92. 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*
93. Atoms are most stable when they have 8 valence electrons (an ***octet***) and tend to form ions to obtain such a configuration of electrons.
Which of the following atoms forms a stable ion that does not have an octet structure?
Li *F* *Na* *Cl*
2- ~ 1s2

94. ***Covalent bonds*** form when two atoms ***share*** a pair of electrons.

95. ***Ionic bonds*** form when one atom ***transfers*** an electron to another atom when forming a bond with it.

Which substance exhibits ionic bonding rather than covalent-bonding?

CO_2 N_2O_4 SiO_2 $CaBr_2$ $C_6H_{12}O_6$

96. **Dot models** may be used to represent the formation of ions or covalent molecules.

. Given the equation:

$:\ddot{F}\cdot + 1e^- \rightarrow [:\ddot{F}:]^-$

This equation represents the formation of a

fluoride ion, which is smaller in radius than a fluorine atom

fluoride ion, which is larger in radius than a fluorine atom

fluorine atom, which is smaller in radius than a fluoride ion

fluorine atom, which is larger is radius than a fluoride ion

97. Substances containing mostly covalent bonds are called ***molecular substances.***

Which of the following is a molecular substance?

Lithium chloride *carbon monoxide* *sodium nitrate* *aluminum oxide*

98. ***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 Hydrogen bonding NOF

HCl *HBr* *HF* *HI*

99. Substances containing mostly ionic bonds are called ***ionic compounds.***

They are made of metal and nonmetallic ions. They are held together by electrostatic (ionic) forces.

100. Liquids ***boil*** when their vapor pressure is equal to the atmospheric pressure. (Reference Table H)

Water will boil at 90°C when the atmospheric pressure is 70 *kPa.*

101. The ***normal boiling point*** of a substance is the temperature at which it boils at 1 atm pressure.

(Reference Table H) @ 101.3kPa = 56°C (55 NOT acceptable)

What is the normal boiling point of propanone?

102. ***Combustion reactions*** occur when a hydrocarbon reacts with oxygen to make CO_2 and H_2O.

103. **Metallic bonds** can be thought of as a crystalline lattice of cations surrounded by a "sea of mobile valence electrons."

Metallic bonding occurs between atoms of...

sulfur *sodium fluoride* *sodium* *carbon*

104. **Non-polar Covalent bonds** form when two atoms of the same element bond together

105. **Polar covalent bonds** form when there is an electronegativity difference greater than) but less than ionic (not a metal and a non-metal) *Which of the following combinations would form a polar covalent bond?*

H and H *Na and N* *H and N* both NM *Na and Br*

106. **Van der Waals** attractive forces are the attractions between non-polar molecules, Non-polar molecules are molecules with structural symmetry.

107. **Intermolecular forces** in non-polar molecules become stronger with increasing molar mass which causes the boiling point to increase. *Which of the following samples has the greatest forces of attraction?*

F_2 Cl_2 Br_2 I_2

108. **Polar molecules:** have stronger forces of attraction and they lack structural symmetry (they are asymmetrical). *Which of the following is a polar molecule?*

CO_2 H_2O C_4H_{10} N_2

109. Memorize this table of properties of the different types of compounds:

Substance Type	Properties
Ionic	Hard *(Low/high) melting and boiling points* Conduct electricity when molten or aqueous
Covalent (Molecular)	Soft *(Low/high) melting and boiling points* Do not conduct electricity (insulators)

Using Table T, solve the following problems:

Problem	Which equation do you need?	Solution (SHOW WORK)
1. If the accepted value for the mass of an object is 10.3g and a student found that the mass was 10.1g, what is the student's percent error?	% error	$\frac{10.1-10.3}{10.3} \times 100$ -1.94%
2. If a peanut is burned in a calorimeter containing 50g of water, and the water temperature changes from 45°C to 57°C, how many joules of energy were released by the peanut?	$Q=mc\Delta T$	$Q=(50)(4.18)(12)$ $=2508J$
3. How much heat does it take to convert 20g of water to steam at 100°C?	$Q=mH_v$	$Q=20(2260)$ $=45200J$
4. A sample of gas has a volume of 60. mL at 40.0 kPa. What will be the new volume of the gas if the pressure is increased to 75.0 kPa and the temperature remains constant?	$\frac{P_1V_1}{T_1}=\frac{P_2V_2}{T_2}$	$(60)(40)=(75)x$ $x=32mL$
5. What is the percent by mass of oxygen in NaOH?	% Composition	$\frac{16}{23+16+1} \times 100$ $\frac{16}{40} \times 100 = 40\%$

22. What is the total number of electrons in the principle energy level that has the greatest amount of energy for an atom of calcium in the ground state?	2-8-8-2 furthest away = most energy therefore 2	Per. Table
23. Which element on the Periodic Table has the strongest attraction for electrons in a chemical bond?	F (4.0 electronegativity) Fluorine	S
24. State the trend in metallic character as you move down Group 14.	metallic character ↑ least electronegative element	S
25. How many kilojoules are equivalent to 10 joules?	$\frac{10}{1000}$ = .01 kJ	C
26. What is the normal boiling point of water in degrees Celsius?	100°C ("normal" = 101.3kPa-find temp)	H
27. Which substance has the strongest IMFs? Your choices are: ethanoic acid, propanone, ethanol, or water.	Ethanoic acid (↑ BP @ a pressure) & (↓ VP @ a temp)	H
28. What is the total number of protons, neutrons, and electrons in the most common isotope of sodium?	mass # 23 P = 11 N = 12 E = 11 equal bcs neutral	per. table

10. What is the electronegativity of chlorine?	3.2	S
11. What is the ionization energy of Rb?	403 KJ/Mol	S
12. Which atom is more likely to lose electrons, Al or Zn?	Al 578 Zn 906 more ionization energy required to remove from Zn	S
13. What is the atomic number of Te?	52	Per table
14. What is the atomic radius of Bromine?	117 pm	S
15. What is the oxidation state of sulfur?	-2 (or +4, +6)	per. table
16. Write the electron configuration of potassium.	2-8-8-1	per. table
17. At what temperature will water boil, when the atmospheric pressure is 55 kPa?	83±1	H
18. What is the trend of atomic radii across period 3?	atomic radius decreases across a period	S
19. What is the heat of vaporization of water?	2260 J/g	B
20. What is the density of tin?	7.287 g/cm³	S
21. What phase of matter is bromine in at STP?	332 Boiling pt g	S

266 melting pt 57 273

Name_______________________________________

Reference Table Scavenger Hunt
Chemistry Midterm Review

Directions: Using the Reference Tables for Chemistry, locate the following information.

Question	Answer	What table did you use?
1. What is the definition of STP, and give the values?	101.3kPa or 1atm 273K 0°C	A
2. Name $C_2H_3O^-_2$ or CH_3COO^-	acetate	E
3. What is the freezing point of fluorine? same as melting	53K	S
4. What are the units for the heat of fusion, and what do they mean?	J/g - heat per mass	D
5. What is the symbol for the mole?	mol	D
6. What is the vapor pressure of water at 75°C?	38 ± 1	H
7. What is the prefix for 1/1000 of a meter?	milli	C
8. What is the formula for the permanganate ion?	MnO_4^{-1}	E
9. What is the atomic mass of silver? What does it represent?	107.868 u weighted average of all naturally occuring isotopes	Per. Table