THIS PACKET IS DUE THURSDAY WHEN YOU TAKE THE

FPS – Semester 2 FINAL Review

EXAM.

Name____________________________ Period__________

1. Define the following terms:
   a. Matter is stuff that has mass and takes up space
   b. Atom – smallest particle of an element that has the same properties of that element.
   c. Pure substance – material comprise of only one type of substance.
   d. Element – substance containing atoms of the same atomic #
   e. Compound – substance made of two or more element combined.
   f. Mixture – matter made of two or more elect
   g. Homogeneous – uniform mixture solution
   h. Heterogeneous – non-uniform mixture suspension/colloid
   i. Solutions – homogeneous mixture
   j. Suspensions – hetergenous mixture where the phases settle out according to density.
   k. Colloids – heterogeneous mixture where particle are so small they do not settle out.
   l. Malleable – could be hammered into different shapes.
   m. Brittle – Easily broken and shattered.
   n. Ductile – Drawn out into long wires.
   o. Conductivity – will allow electricity to flow through
   p. Melting point – temp at which solid turns into a liquid
   q. Boiling point – temp at which a liquid is turned into a gas.
   r. Density – measure of how heavy something is relative to area.
   s. Viscosity – the measure of a liquid ability to flow.
   t. Flammability – ability for a substance to burn
   u. Reactivity – the ability of a substance interact with its environment
   v. Boyle's Law as pressure increases vol decreases.
   w. Charles's Law as temperature increase volume increases,
   x. Immiscible – Cannot Mix
   y. Temperature – Av KE of the particles in a substance.
   z. Pressure – force exerted per unit area.

2. List several physical properties.

   b.p., m.p., density, temp, color, etc.

3. How can physical properties be used?

IDentify Substances

4. List two chemical properties.

Burning, Flammability.

5. What is the different between a physical change and a chemical change?

Chemical bonds are broken.

6. Give 3 examples of a physical change and justify.

Phase Change, Dissolving,

7. Give 3 examples of indications of a chemical change. Then, give 3 examples of different chemical changes.

Color change Solid Produced. Fire
Gas produced Energy exchange

8. What is the Tyndall effect and which mixture does it identify?

Tyndall effect is caused by the scattering of light as it ~~passed~~ through a liquid (Suspension + colloids)

9. Give several examples of elements.

Hydrogen, Carbon
Helium, Sodium

10. Give several examples of compound.

NaCl Magnesium Sulphate
Glucose CO_2

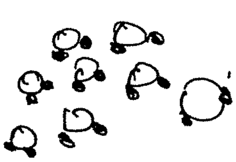
11. How is a homogeneous mixture different from a heterogeneous mixture? Give several examples of both.

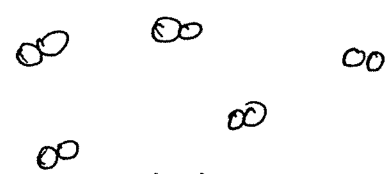
Totally uniform

Separate into phases according to density.

12. List and describe the three states of matter and SKETCH the particle arrangements.

Solid





very compact, strong attractive forces
it vibration motion
very little energy

liquid -
particle have more motion
weaker attractive force
moderate energy.

gas.
particles have total random constant
almost zero attractive fo

13. Which state of matter has the highest energy and why?

Gas has the highest because it has the weakest forces of attraction and move the fastest

14. List 5 phase changes and define them.

Melting Condensation
Freezing Sublimation.
Boiling

15. Why is a phase change a physical change?

No chemical bonds are broken.

16. Give 2 examples that demonstrate Boyle's Law.

Squeeze (↓ volume) a balloon → ↑ pressure
Place a balloon in a vacuum (↓ pressure) → ↑ volume

17. Give 2 examples that demonstrate Charles's law.

Exploding cans with ↑ heat
Flat tire in winter

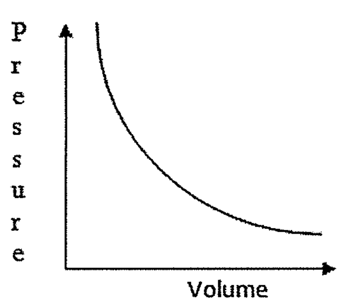
18. Explain why bicycle tires seem more flat in the winter.

↓ temp → ↓ pressure
↳ ↓ volume

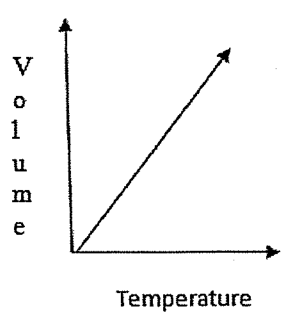
19. Explain why a can of soda explodes if left in the sun.

↑ Temp → ↑ pressure
↳ ↑ volume

20. Steel has a density of 7.8 g/cm3. What is the mass of the block of steel with a volume of 600cm3? (Hint 1mL = 1cm3) $D = \frac{m}{V}$ so $m = DV = 7.8g/\cancel{cm^3} \times 600\cancel{cm^3} = \boxed{4680\ g}$
21. A substance has a mass of 360 g and a volume of 7.5 cm3. What is its density? $D = m/V = \frac{360g}{7.5\ cm^3} = \boxed{48 g/cm^3}$
22. Identify each of the following as a compound or an element.
    a. Cl ____E____
    b. CH_4 ____C____
    c. Co ____E____
    d. CO_2 ____C____
23. State the laws that fit with the following graphs.

a. 

Boyles

b. 

Charles

A. History of Atomic Theory

24. Summarize the following people's discoveries and ideas. Dates/drawings will not be required on the test.
    a. Democritus – atmos
    b. *Atomos* – indivisible particles.
    c. Dalton – atom
    d. Thomson – electron
    e. Plum-pudding model – Thompson's Experimental model
    f. Rutherford – nucleus + protons.
    g. Gold-foil experiment – Rutherford's experiment.
    h. Bohr's model – planetary model.
    i. Electron-cloud model – Schrodinger and Heisenberg and DeBrolige

25. What were some differences between Dalton and Democritus's ideas?

includes compounds

26. Who described the Billiard ball model? Why was it called such?

solid piece of indivisible matter

27. Rutherford, who was Thomson's student, refuted the plum-pudding model. Describe how his experiment did this.

empty space — solid sphere with charge dispersed

positive nucleus w/ electrons on the outside.

28. How is Bohr's model different than previous models?

Electron are found in specified orbitals.

29. How is the electron-cloud model different than previous models? Specific location of orbital is only a probability

30. What were Dalton's three parts of his Atomic Theory of matter?

All atoms of an element are the same

Atoms are small and indivisible

Compounds combine atoms in a fixed ratio.

31. According to Bohr's model of the atom, where are the electrons? What can happen for them to change location? In circular orbitals circling the nucleus.

Absorbe energy and be tranferred to a higher energy level.

B. The Periodic Table & Atomic Structure

32. How is the modern periodic table of element arranged?

By atomic #

33. Assuming the atom is neutral, what does the atomic number tell you?

protons | # neutrons.

34. Label the following periodic table square for argon.

Argon	Name
18	Atomic #
Ar	Symbol
39.948	Atomic Mass.

35. What is a period? What does a period on the periodic table indicate about an element?

Row — # energy levels.

36. What is a group/family? What does a group/family on the periodic table indicate about an element?

Column — # of valence electrons.

37. What do valence electrons indicate about an element?

Group #

38. What is an ion? What do we call a positive and negative ion?

atoms that have lost or gained electrons — cation — anion.

39. What is the octet rule? What are the exceptions?

Atoms aspire to get 8 valence electrons — H + He

40. What does the mass number tell you? What are isotopes?

protons + neutron — atoms with the same # proton by different # neutrons.

41. In nuclear notation, write the isotopes magnesium-24, magnesium-25, and magnesium-26.

$^{24}_{12}Mg$ $^{25}_{12}Mg$ $^{26}_{12}Mg$

42. What is the average atomic mass of Boron if it exists as 19.90% ^{10}B and 80.10% ^{11}B?

$$(10 \times 0.1990) + (11 \times 0.8010) = 10.801 \text{ amu}$$

43. Magnesium has three naturally occurring isotopes. 78.70% of Magnesium ~~atoms exist as~~ Magnesium-24, 10.03% exist as Magnesium-25 and 11.17% exist as Magnesium-26. What is the average atomic mass of Magnesium?

$$(24 \times 0.7870) + (25 \times 0.1003) + (26 \times 0.1117) = 24.30 \text{ amu}$$

C. *Element Categories*

44. Describe properties of the alkali metals and give an example.

Very reactive w/. water + air — Na

45. Describe properties of the alkaline-earth metals and give an example.

Denser than water Less reactive in water + air — Ca

46. Describe properties of the transition metals and give an example.

Magnetic Fe — Colored Gold

47. Describe properties of the metalloids and give an example.

Conduct at high temps — Si

48. Describe properties of the halogens and give an example.

Very reactive with Alkali metals — Cl, Br, I, F.

49. Describe properties of the noble gases and give an example.

Very unreactive due to complete octet, He, Ne, Ar, Kr, Xe, Rn

50. What happens to metallic properties as you move from left to right on the periodic table?

Become more non-metallic.

51. What happens to reactivity as you down a group on the periodic table?

Metal become — Non-metal

52. WHY do the alkali metals become more reactive down the group?

Valence electron further from nucleus easier to react.

53. WHY do the halogens become more reactive up the group?

Valence electrons closer to the nucleus easier to attract an electron.

D. *Electron Configurations*

54. What are the 4 types of orbitals? How many electrons can each of them hold?

s, p, d, f
2 6 10 14

55. Write the complete electron configuration for the following elements:

a. Potassium $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
b. Lithium $1s^2 2s^1$
c. Aluminum $1s^2 2s^2 2p^6 3s^2 3p^1$
d. Carbon $1s^2 2s^2 2p^2$
e. Nitrogen $1s^2 2s^2 2p^3$
f. Argon $1s^2 2s^2 2p^6 3s^2 3p^6$

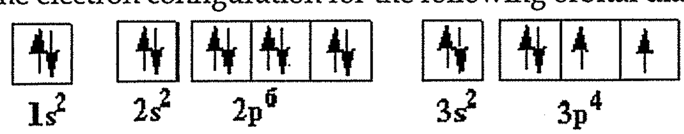
56. Write the electron configurations for the following *ions.* Remember, if an ion is positive, it has lost electrons. If an ion is negative, it has gained electrons.

a. Be^{2+} $1s^2$
b. B^{3+} $1s^2$
c. Cl^{1-} $1s^2 2s^2 2p^6 3s^2 3p^6$
d. O^{2-} $1s^2 2s^2 2p^6$

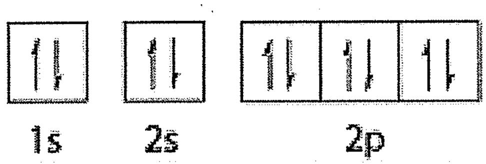
57. Identify the errors in the following electron configurations. If there is no error, write "none".

a. $1s^2 2s^3 2p^2$ None
b. $1s^2 2s^2 2p^6 3s^2 4s^1$ Should be $3p^1$
c. $1s^2 2s^2 2p^8 3s^1$ $2p^6$ only.

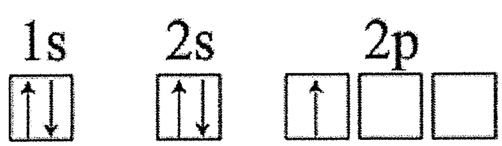
58. Write the electron configuration for the following orbital diagrams. Then, identify the element.

S


a.

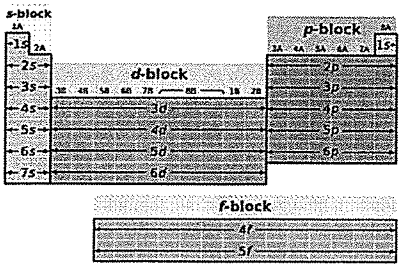
Ne


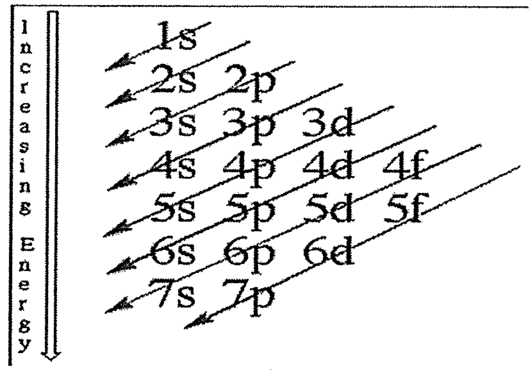
b.

B


c.

Use your periodic table and the images and formulas below as reference.





Number of Neutrons + Number of Protons = Mass Number

$$average\ atomic\ mass = \left(isotope\ mass\ x\ \frac{\%}{100}\right) + \left(isotope\ mass\ x\ \frac{\%}{100}\right) + \cdots$$

E. *Bonding*

59. Define the following bonding types and give examples of the compounds that are bonded in that way.
   a. Ionic Force of attraction between cation + anion
      i. Examples: NaCl.
   b. Covalent Force of attraction between the nucleus of a neutral atom and its neighbors electrons.
      i. Examples: CO_2
   c. Metallic Force of attraction between metal cation and ~~[illegible]~~ dissociate valence electrons.
      i. Examples: Fe

60. What are some identifying properties of ionic bonds? How are the ions arranged?
High melting + boiling points 3-D Lattice Crystal.
Conducts electricity when liquid or dissolved in water.

61. What are some identifying properties of covalent bonds? How are the units different from ions?
Mostly Low melting and boiling points → Made of neutral atoms.
Don't conduct electricity when molten only acids do when dissolv

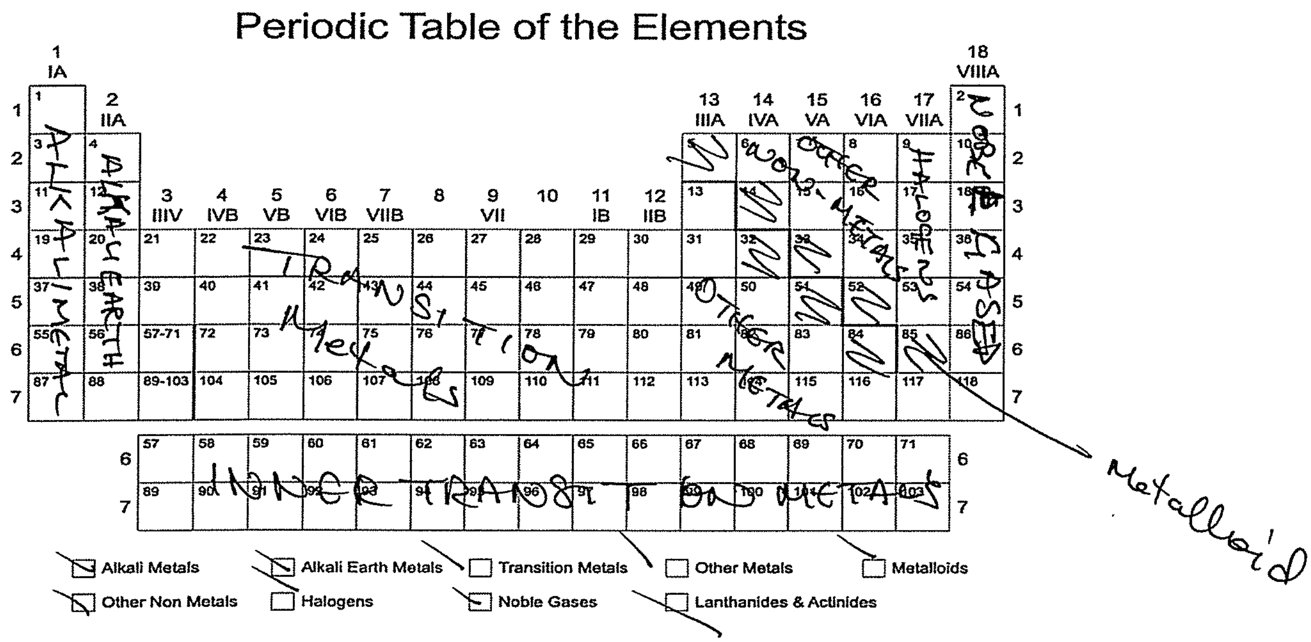
62. What are some identifying properties of metallic bonds? How are the electron interactions unique?

63. Describe the differences in valence electron behavior for each of the 3 bonding types.

64. What is the octet rule? How do atoms satisfy the octet rule?
Atoms aspire for 8 Valence Electron → gain / lose or share valence electrons.

65. How are the bonds in H_2O different from the bonds in Li_2O?
↓ covalent ↓ ionic

66. On the periodic table below, identify the 3 *main categories* of elements. Then, label the charges for each column.



F. ***Lewis Dot Structures and Bonding***

67. What is depicted in a Lewis dot structure?

Element Symbol and valence electrons.

68. Draw the Lewis dot structures for the following elements:
   a. Carbon ·O·
   b. Aluminum ·Al·
   c. Lithium Li ·
   d. Helium ·He ·
   e. Argon :Ar:
   f. Phosphorus P
   g. Silicon S
   h. Oxygen :O:
   i. Fluorine :F·

69. Draw the Lewis dot structures for the following compounds. Remember, subscripts tell you the number of atoms present per unit.
   a. H_2O :O — H, H
   b. CH_4 H—C—H, H, H
   c. LiCl [Li:]$^+$ [:Cl:]$^-$
   d. Na_2S 2 [Na]$^+$ [:S:]$^{2-}$
   e. O_2 :O = O

70. Write the chemical formulas for the ionic compounds containing the following elements. Refer to #8 for charges to use the crossing method.

a. Na and O Na_2O

b. Ca and Cl $CaCl_2$

c. Mg and N Mg_3N_2

d. Cs and F CsF

C. Naming Binary Compounds

71. What are the differences in naming covalent and ionic compounds?

Greek prefixes

72. When do you NOT use a prefix for a covalent compound?

when first elect have mono.

73. Fill in the prefix table below.

Prefix	*Number*
Mono	1
di	2
tri	3
tetra	4
penta	5
hexa	6
hepta	7
octa	8
nona	9
deca.	10

74. Name the following *ionic* compounds:

a. NaCl Sodium Chloride

b. MgF_2 Magnesium Fluoride

c. $AlCl_3$ Aluminium Chloride

d. BeO Beryllium Oxide

75. From the following names, write the chemical formula for the ionic compounds. Write the ions first, then use the crossing method.

a. Sodium oxide Na_2O

b. Magnesium oxide MgO

c. Barium fluoride BaF_2

d. Lithium bromide LiBr

76. Write the names for the following *covalent* compounds.

a. CO carbon monoxide

b. CO_2 carbon dioxide

c. N_2H_4 dinitrogen tetra hydro

d. SO_4 sulfur tetraxide

e. N_3O_5 trinitrogen pentoxide

f. CS_6 carbon hexasulfide

77. Write the formula for the following names.

a. trisilicon tetrafluoride

Si

b. carbon trioxide CO_3
c. dichlorine heptoxide Cl_2O_7
d. tetracarbon decasulfide C_4S_{10}
e. boron hexachloride BCl_{10}
f. dihydrogen dioxide H_2O_2

78. Below is a mixed set of chemical compounds. Ionic and covalent are both present. Name or give the chemical formula.

a. CaO Calcium Oxide
b. H_2O_2 dihydrogen dioxide
c. Carbon dioxide CO_2
d. Magnesium oxide MgO
e. PCl_3 ~~phosphorus~~ Phosphorous Trichloride
f. NH_2 nitrogen dihydrogen
g. Dinitrogen monoxide N_2O
h. Potassium bromide KBr
i. Nitrogen trioxide NO_3
j. NO nitrogen monoxide
k. SF_6 sulfur hexafloride
l. Strontium nitride Sr_3N_2
m. Diphosphorus pentoxide P_2O_5

G. Definitions

79. Define and give the "formulaic" pattern for each type of reaction.

a. Synthesis

$A + B \rightarrow AB$

b. Decomposition

$AB \rightarrow A + B$

c. Single displacement

$A + BC \rightarrow AC + B$

d. Double displacement

$AB + CD \rightarrow AD + CB$

e. Combustion

$C_xH_y + O_2 \rightarrow CO_2 + H_2O$

80. Describe many observations that can be made when a chemical reaction is occurring.

Color change
Production of solid
Prod. gas.

81. What is a precipitate?

Solid produced in a reaction

82. List AND define the parts of a chemical equation.

reactants — catalyst ~~heat~~ pressure / yield arrow — Products

83. Define the Law of conservation of mass.

Mass cannot be created or destroyed only changed ~~[illegible]~~ and rearrange chemical bonds.

84. Why do we balance chemical equations?

to obey the law of conservation of mass

85. Why can't you change the subscripts of a chemical equation when balancing?

You'll have different chemicals than the observed schedule

86. What is the rate of a reaction? What are the four factors that can affect the rate of a reaction?

How fast products are made or reactants are consumed.

Surface area
Temperature
Concentration of reactants
Enzyme/Catalysts.

87. What is activation energy?

Energy required to kick start a reaction.

88. Why doesn't a reaction last forever and constantly increase its rate?

Eventually all the reactants are consumed

89. How and WHY does temperature affect the rate of reaction?

Speeds up the motion of the reactants. This allows for a higher frequency of collision and an increased energy of collision.

90. How and WHY does increased surface area speed up the rate of reaction?

It allows for more molecules to collide.

91. How and WHY does increased concentration speed up the rate of reaction?

There are less collisions with particles not involved in the reaction.

92. How and WHY does using a catalyst speed up the rate of reaction?

Provides an active site that allows for collisions to occur in the correct orientation.

93. What three catalysts are commonly used in industry? Why are catalysts important?

They speed up the reaction without being consumed

94. What is the difference between endothermic and exothermic reactions? How can we tell in the lab?

Exo ↑ temp
Endo ↓ temp

absorb more heat than the system gives off

gives off more heat than the system absorbs.

95. What is a molar mass?

Mass of 1 mole of a substance in grams.

96. How do you calculate a molar mass? Explain using NaCl as an example.

Take the atomic weight of each element in the substance, multiply each by the subscript in the formula and then add up the products.

H. *Application*

Balance and write the type of reaction.

		Type of reaction:
97.	$4Na + O_2 \rightarrow 2Na_2O$	Synthesis
98.	$SF6 \rightarrow S + 3F_2$	Decomposition
99.	$2Al + 3CuSO_4 \rightarrow 3Cu + Al_2(SO_4)_3$	single replacement
100.	$Pb(NO_3)_2 + 2NH_4Cl \rightarrow PbCl_2 + 2NH_4NO_3$	Double replacement
101.	$Al_2O_3 + 3H_2O \rightarrow 2Al(OH)_3$	Synthesis
102.	$2Fe_2O_3 \rightarrow 4Fe + 3O_2$	Decomposition
103.	$Zn + SnCl_2 \rightarrow Sn + ZnCl_2$	single replacement
104.	$Cu(NO_3)_2 + Na_2S \rightarrow CuS + 2NaNO_3$	double replacement
105.	$2Al(OH)_3 + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 6H_2O$	double replacement
106.	$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$	Combustion
107.	$2Ag + S \rightarrow Ag_2S$	synthesis
108.	$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3$	single replacement
109.	$2HgO \rightarrow 2Hg + O_2$	decomposition
110.	$KCl + AgNO_3 \rightarrow AgCl + KNO_3$	double replacement
111.	$Al_2(SO_4)_3 + 3Ca(OH)_2 \rightarrow 3CaSO_4 + 2Al(OH)_3$	double replacement
112.	$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3$	single replacement
113.	$4Al + 3O_2 \rightarrow 2Al_2O_3$	synthesis
114.	$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$	combustion

115. Draw an energy diagram for an endothermic reaction.

116. Draw an energy diagram for an exothermic reaction.

117. *Calculate the molar masses of the following chemicals:*

1) Cl_2 $2(35.5)$ 71 g/mol

2) KOH $39 + 16 + 1$ 56 g/mol

3) $BeCl_2$ $9 + 2(35.5)$ 80 g/mol

4) $FeCl_3$ $56 + 3(35.5)$ 162.5 g/mol

5) BF_3 $11 + 3(19)$ 68 g/mol

6) CCl_2F_2 $12 + 2(35.5) + 2(19)$ 121 g/mol

7) $Mg(OH)_2$ $24 + 2(16) + 2(1)$ 58 g/mol

8) UF_6 $238 + 6(19)$ 352 g/mol

9) SO_2 $32 + 2(16)$ 64 g/mol

10) H_3PO_4 $3(1) + 31 + 4(16)$ 98 g/mol

11) $(NH_4)_2SO_4$ $2(14) + 8(1) + 32 + 4(16)$ 132 g/mol

12) CH_3COOH $2(12) + 4(1) + 2(16)$ 60 g/mol

13) $Pb(NO_3)_2$ $207 + 2(14) + 6(0)$ 331 g/mol

14) $Ga_2(SO_3)_3$ $2(70) + 3(32) + 9(16)$ 380 g/mol

117. Describe **homogeneous mixtures.** uniform mixture

118. Describe **heterogeneous mixtures.** non-uniform mixture

119. Describe the parts of solutions (solute and solvent).

less more

120. Describe a suspension.

Heterogeneous mixture who's particle settle out according to density.

121. Describe a colloid.

Heterogeneous mixture who's particle are small enough that they do not settle out.

122. Describe the terms unsaturated, saturated, and supersaturated.

can dissolve more — cannot dissolve more — made to dissolve more than the solubility allows.

123. In the lab, how did we *supersaturate* a solution?

Saturate heat and add more than allow to cool.

124. Describe the ways we can alter solubility.

Heat, Pressure

125. What is a base? Give some examples. NaOH, KOH

Solution that increase OH^- concentration is solution.

126. Give the properties of a base. Slippery to the touch. Bitter, Corrosive, pH greater than 7, will neutralize an acid

127. What is an acid? Give some examples.

solution that increased H^+ concentrat

128. Give the properties of an acid. wet to the touch. Sour, Corrosive, pH less than 7, will neutralize a bas

129. Which ion does a base give off? An acid?

OH^- H^+/H_3O^+

130. Why do acids and bases conduct electricity?

They have free moving ion in solution.

131. When an acid and a base come together, what happens?

They neutralize each other to give salt + water.

132. What colors do Acids and Bases turn litmus paper?

Blue -> Red Red -> Blue.

133. Describe the pH scale. What does it measure? Where do acids fall? Bases?

logorithmic scale from 0 - 14 | Acidity | < 7 = pH

134. Using what you know about the litmus colors produced by acids and bases, fill in the table below.

Red litmus paper	*Blue litmus paper*	*pH*	*Acid, base, or neutral?*
Red	Blue	7	neutral
Blue	Blue	11	Base
Red	Red	4	acid
Blue	Blue	8.5	base
Red	Blue	7	Neutral
Red	Red	1	Acid

135. If you test two fruits, and find that the lemons has a pH of 3 and the apple a pH of 5, which is the **stronger** acid? Why?

Lemon (↑ [H+]) higher hydrogen concentration.

136. What is the nucleus of an atom composed of?

Proton + neutrons.

137. What is a strong force?

138. How are the strong forces between a large nucleus and a small nucleus different?

139. Why does nuclear decay happen?

For a nucleus to become more stable.

140. What two things are released during nuclear decay?

Particle (α, β) and Energy (γ)

141. What is an isotope?

Same atomic # different mass #

142. Give an example of isotopes of Carbon.

$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$

143. What does it mean when a nucleus is unstable or *radioactive?*

# of proton to neutron unfavorable → 'unstability'

144. In what elements did Marie Curie discover radioactivity?

Radium Polonium Uranium.

145. What is an alpha particle?

$^{4}_{2}He$

146. What is a beta particle?

electron $^{0}_{-1}e$

147. What is a gamma ray?

Electromagnetic Radiation

148. How are alpha and beta particles different?

large slow positive

tiny fast negative

149. How are gamma rays different from alpha and beta particles?

wave/photons

particle or matter

150. Which is the most dangerous radiation AND why?

γ highest energy can penetrate the deepest.

151. Complete the following equations.

a) $^{14}_{6}C \rightarrow \ ^{0}_{-1}e +$ $^{14}_{7}N$

b) $^{229}_{90}Th \rightarrow \ ^{4}_{2}He +$ $^{225}_{88}Ra$

c) $^{20}_{8}O \rightarrow \ ^{20}_{9}F +$ $^{0}_{-1}e$

d) $^{238}_{92}U \rightarrow \ ^{234}_{90}Th +$ $^{4}_{2}He$

e) $^{207}_{81}Tl$ $\rightarrow \ ^{207}_{82}Pb + \ ^{0}_{-1}e$

152. List the type of radiation shown in a-e in question one.

a) Beta

b) Alpha

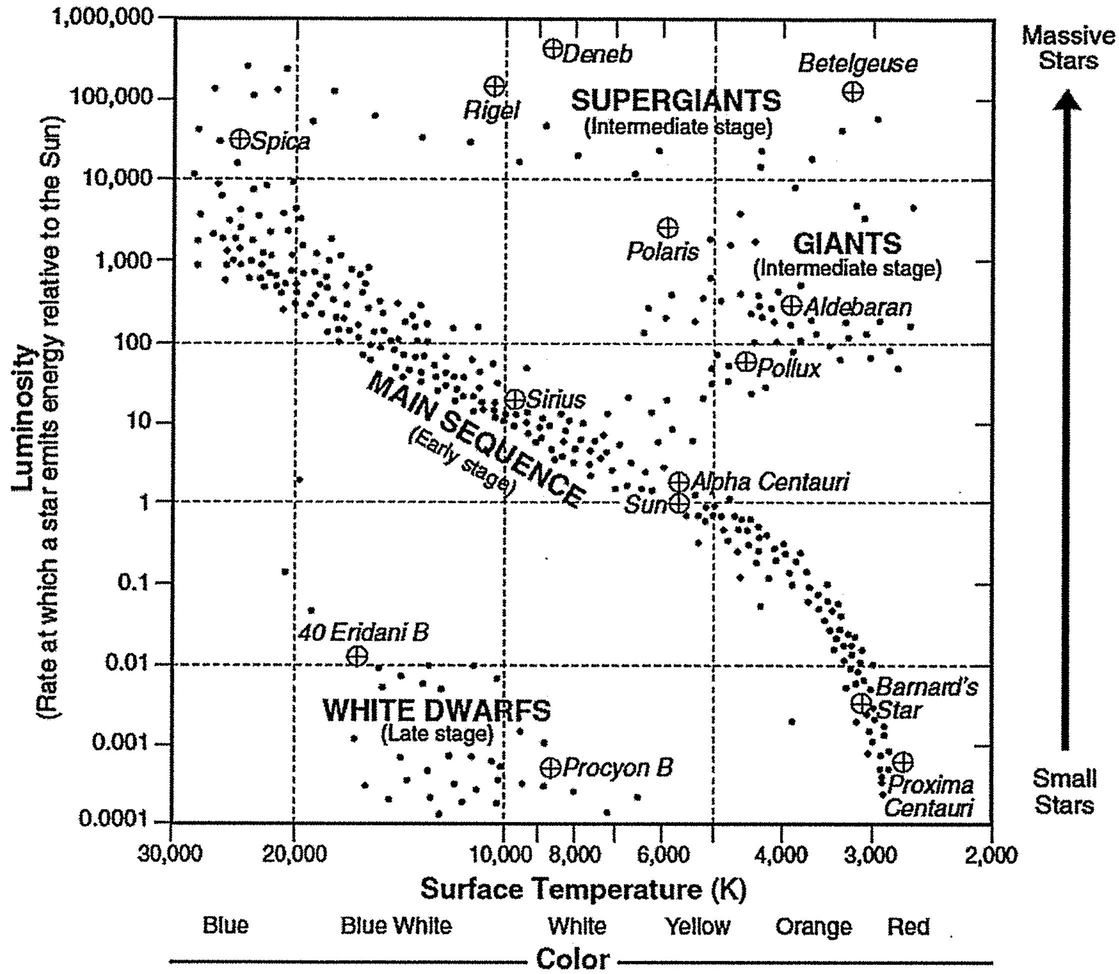
c) Beta

d) Alpha

e) Beta.

Characteristics of Stars

(Name in italics refers to star represented by a ⊕.)
(Stages indicate the general sequence of star development.)



1. Describe the size of stars in the H-R diagram (in comparison with the sun) in the upper right portion:

much larger somewhat larger the same size somewhat smaller much smaller

2. Describe the color of the stars in the upper right portion of the H-R diagram.

redder about the same color bluer

3. Describe the color of the stars in the upper left portion of the H-R diagram.

redder about the same color bluer

4. Where in the H-R diagram would you find a star with a greater temperature than the sun?

left right above below

5. Where in the H-R diagram would you find a star with a greater luminosity than the sun?

left right above below

6. What is the classification of the Sun? Main Sequence

7. What is the classification of Rigel? Blue Super Giant

8. Name a star that is a Blue. Spica

9. Which star is most similar to the sun? Alpha Centauri

10. What is the temperature of the sun? ~ 6000 K

11. Name a star that is hotter than the sun but smaller in size 40 Eridani B

12. Name a star that is cooler than the sun but larger in size Pollux

13. How much dimmer than the sun is Eridani B? X 100

14. How would you characterize the stars that are both very bright and cooler?

blue giants   red giants   white dwarfs   red dwarfs   sun-like

15. Name on star that is brighter than Aldebaran RIGEL

16. What is the hottest main sequence star? SPICA