CHEMISTRY SOL REVIEW MATERIAL Name_______________________

SCIENTIFIC INVESTIGATION

1. On previous SOL tests, students have been asked to choose the piece of glassware that gives the most *precise* results. They have usually been given the following choices:

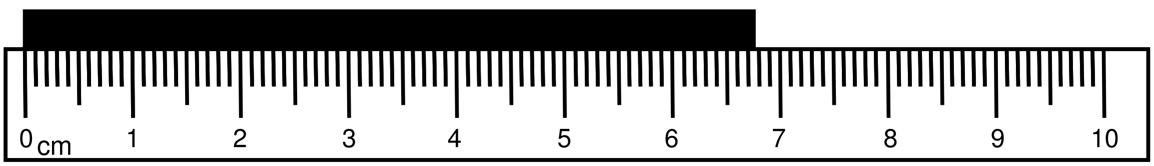
   beaker   flask   pipet   test tube   graduated cylinder

   Of these choices, the most precise piece of glassware is the **GRADUATED CYLINDER**

   because **EACH LINE REPRESENTS 1 mL or 0.1 mL. SMALL DIVISIONS BETWEEN LINES = MORE PRECISE**

2. When you read the volume on a graduated cylinder or you measure length with a ruler, you should always estimate the final digit. The final estimated digit will always be one power of ten *smaller* what each line is worth on the instrument. In other words, if the ruler shows lines every 0.1 cm, then you estimate length to the nearest 0.01 cm.

   Estimate the length of this strip: **6.75** cm



3. If you measure something in an experiment, why do you think it is a good idea to perform the measurement three separate times and take the average result?

   TO VERIFY YOUR RESULTS; TO GET CONSISTENT, REPRODUCIBLE RESULTS

4. If you take several measurements, then your data will be *precise* if

   THE NUMBERS ARE VERY CLOSE TO EACH OTHER

5. Data is considered to be *accurate* if **THE NUMBERS ARE CLOSE TO THE ACCEPTED VALUE**

6. A common scenario is to show data that is *precise but not accurate*. The boiling point of water is 100.0°C. Give an example of data for the BP of H_2O that is *precise but not accurate*:

   Trial 1: **81°C**   Trial 2: **82°C**   Trial 3: **82°C**   Trial 4: **81°C**   (answers may vary)

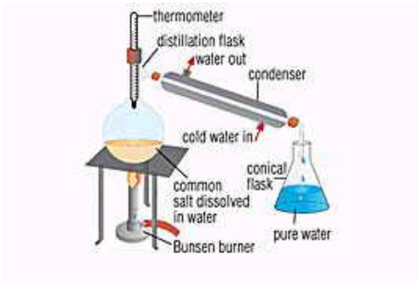
6. Basic lab techniques for separation of a mixture are listed below. Match the physical property with the separation technique.

   C chromatography   A. boiling point

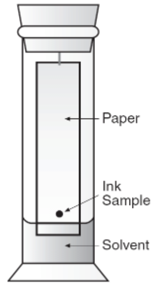
   B filtration   B. particle size

   A distillation   C. interaction with the solvent (polarity)

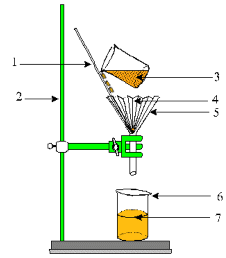
7. Now match the separation technique with the picture:



A.



B.



C.

___B___ chromatography

___C___ filtration

___A___ distillation

8. Fill in the blanks by writing numbers in either regular notation or scientific notation.

Regular Notation	Scientific Notation
15600	1.56×10^{4}
250,000	**2.5×10^{5}**
0.00045	**4.5×10^{-4}**
2300	2.3×10^{3}
0.0061	6.1×10^{-3}

9. If you get a chemical on your skin or in your eyes, the first thing you should do is always

FLUSH OR RINSE WITH LOTS OF WATER

10. If you need to mix acid and water together, remember that the safety rules state that

you should always add ___**ACID**___ to ___**WATER**___

(Heat may be given off when acids are mixed with water. Concentrated acids are more dense than water; they sink to bottom of water and mix more evenly. Otherwise adding water to acid may cause spattering on top of liquid surface)

11. If you see a graph that shows a relationship between two variables, it will often fall into one of two categories: *direct relationship* or *inverse relationship*.

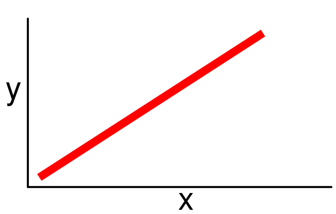
A *direct relationship* can be summarized by saying that as one variable increases, the other

variable ___**INCREASES**___. An example of this would be ___**VOLUME AND TEMP.**___

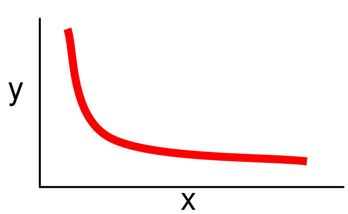
An *inverse relationship* can be summarized by saying that as one variable increases, the other

variable ___**DECREASES**___. An example of this would be ___**PRESSURE AND VOLUME**___

Sketch the general shape of each graph below:



Direct relationship



Inverse relationship

12. If you are asked to calculate *percent error*, you should know that

$$\text{percent error} = \frac{|\text{ measured value} - \text{accepted value }|}{\text{accepted value}} \times 100\%$$

A certain piece of metal has an accepted mass of 65.0 grams. Its mass was recorded in the laboratory as 55.0 grams. Calculate the *percent error* in this measurement.

% ERROR = $\dfrac{\textbf{10.0 g}}{\textbf{65.0 g}}$ x 100% = 15.4%

13. The following information concerns the metric system and other unit conversions. You should definitely know these numbers.

1 L = __**1000**__ mL   1 kg = __**1000**__ g ^{o}C + __**273**__ = K

14. Students often forget how to determine how many *significant figures* are in a given measurement. See if you remember how to do this.

Number	Significant Figures
25.7	**THREE**
100.62	**FIVE**
5.00	**THREE**
200	**ONE**
200.0	**FOUR**
0.075	**TWO**
0.0050	**TWO**

15. When you multiply or divide two numbers, the rule is that the final answer should be rounded so that it has the same number of sig figs as the measurement with the fewest sig figs.

If the mass of an object is 2.7 g and the volume is 3.5 mL, calculate the *density* and round your answer to the proper number of sig figs.

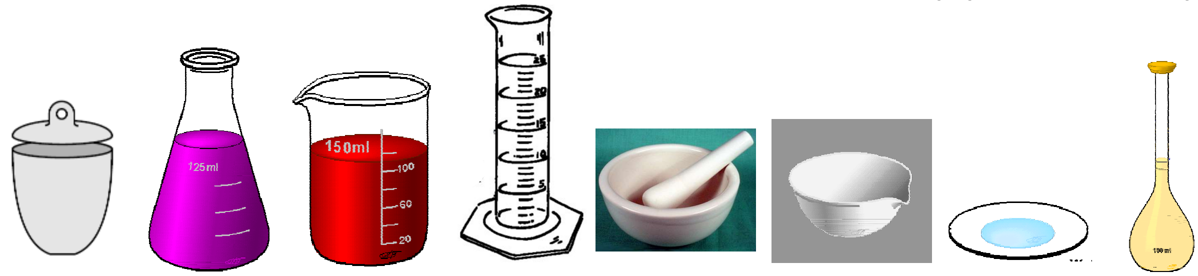
DENSITY = $\dfrac{\textbf{2.7 g}}{\textbf{3.5 mL}}$ = 0.7714 g/mL (unrounded) → 0.77 g/mL (rounded)

16. Report the *average* of these three measurements using the correct number of significant figures.

Trial 1	Trial 2	Trial 3	Average
85.2	84.9	85.4	**85.2**

(unrounded = 85.1666)

17. Sometimes students are asked to identify pieces of laboratory equipment.
In each blank below write the letter that matches the name of the equipment with the picture.



A. B. C. D. E. F. G. H.

___C___ beaker	___F___ evaporating dish	___H___ volumetric flask
___A___ crucible	___D___ graduated cylinder	___G___ watch glass
___B___ Erlenmeyer flask	___E___ mortar and pestle	

ATOMIC STRUCTURE AND PERIODIC RELATIONSHIPS

1. Here are some scientists you should know.

Mendeleev Rutherford Dalton Bohr

___DALTON___ He came up with an atomic theory in 1803 that said that atoms were indivisible building blocks of matter. He thought that all atoms of a given m element were identical.

___RUTHERFORD___ He did a famous gold foil experiment that led him to conclude that all atoms contain a tiny dense center of positive charge called the nucleus.

___BOHR___ He tried to explain the bright-line spectrum of hydrogen with a model of the atom in which electrons occupy fixed energy levels and circle the nucleus in orbits, like planets around the sun.

___MENDELEEV___ He came up with the first periodic table and predicted the properties of a few elements that had not been discovered yet.

2. Elements contain three subatomic particles. Fill in the missing data:

Particle	Charge	Mass Number	Location
PROTON	+	1	**in the nucleus**
NEUTRON	**0**	**1**	**in the nucleus**
ELECTRON	**-**	0	around the nucleus

3. Remember that the *atomic number* refers to the number of ___PROTONS___ in an atom.

The *mass number* refers to the sum of the ___PROTONS___ and ___NEUTRONS___ in an atom.

You should know that atoms are neutral. They have no charge, because they have the same number of ___PROTONS___ and ___ELECTRONS___

4 Fill in the missing information in the table.

Make sure that you got the symbols correct →

Symbol	Protons	Neutrons	Electrons
${}^{23}_{11}Na^{+}$	11	12	10
${}^{31}_{15}P$	15	16	15
${}^{40}_{20}Ca^{+2}$	20	20	18
${}^{80}_{35}Br^{-}$	35	45	36
${}^{39}_{19}K^{+}$	**19**	**20**	**18**
${}^{75}_{33}As^{-3}$	**33**	**42**	**36**

5. When a neutral atom gains or loses electrons, it becomes an ___**ION**__________.

6. An atom that *loses* electrons will have a ___**+**___ charge. This is called a ___**CATION**______

7. An atom that *gains* electrons will have a ___**-**___ charge. This is called an ___**ANION**______

8. If two atoms have the same number of protons, but different numbers of neutrons, these atoms would represent different ___**ISOTOPES**__________of the same element.

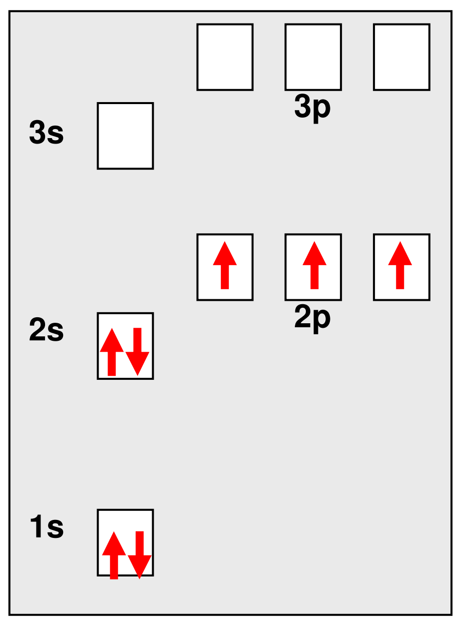
Electrons will fill energy levels according to certain rules. At right is an energy level diagram. Here are the rules:

<u>aufbau rule</u>: start at the bottom and work your way up

<u>Pauli exclusion principle</u>: no more than two electrons in each orbital. Two electrons in same orbital have opposite spins

<u>Hund's rule</u>: when a sublevel has more than one orbital (like the p sublevel) you should always put electrons one at a time into each orbital before you double them up

9. Fill in the electrons in the diagram at the right for the atom NITROGEN. →



10. Fill in the missing information in the table:

Element name	Element Symbol	Complete Electron Configuration	Noble Gas Abbreviated Electron Configuration
magnesium	Mg	$1s^22s^22p^63s^2$	[Ne] $3s^2$
sulfur	**S**	**$1s^22s^22p^63s^23p^4$**	**[Ne] $3s^23p^4$**
calcium	Ca	**$1s^22s^22p^63s^23p^64s^2$**	**[Ar] $4s^2$**
gallium	**Ga**	$1s^22s^22p^63s^23p^64s^23d^{10}4p^1$	**[Ar] $4s^23d^{10}4p^1$**
silicon	**Si**	**$1s^22s^22p^63s^23p^2$**	[Ne] $3s^23p^2$

11. You might see an abbreviated electron configuration notation that looks like this. Fill in the Group number for each electron configuration.

	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
Group	**1**	**2**	**13**	**14**	**15**	**16**	**17**	**18**

12. There are several things about the periodic table that you should know:

Horizontal rows of the periodic table are called **PERIODS**

Vertical columns of the periodic table are called **GROUPS** or **FAMILIES**

Two elements that are located in the same group will have the same number of

VALENCE electrons, and they will have similar **CHEMICAL PROPERTIES**

13. There are seven elements that are diatomic, because they exist naturally in the form X_2. Write the symbols for these seven diatomic elements:

H_2 **O_2** **N_2** **F_2** **Cl_2** **Br_2** **I_2**

14. The following groups or sections of the periodic table have names that you should know. Name them.

Group 1 **ALKALI METALS**

Group 2 **ALKALINE EARTH METALS**

Groups 3-12 **TRANSITION METALS**

Group 17 **HALOGENS**

Group 18 **NOBLE GASES**

15. *Horizontal* trends of the periodic table:

As you move *from left to right across a period* of the periodic table,

THE NUMBER OF PROTONS will __INCREASE_______________

THE ATOMIC RADIUS tends to ____DECREASE_______________

THE 1st IONIZATION ENERGY tends to ____INCREASE________

THE ELECTRONEGATIVITY tends to __INCREASE_____________

16. *Vertical* trends of the periodic table:

As you move *from top to bottom down a group* of the periodic table,

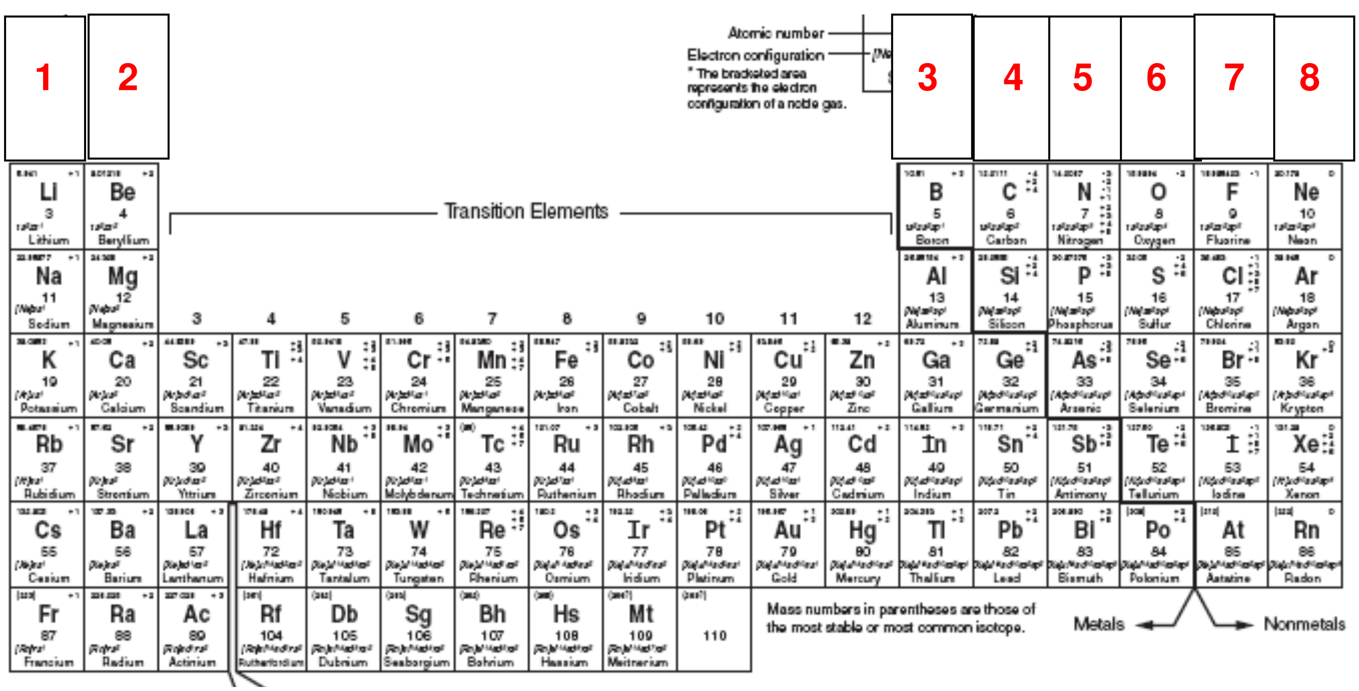
THE PRINCIPAL QUANTUM NUMBER (energy level *n*) will _INCREASE________

THE ATOMIC RADIUS tends to ____INCREASE_______________

THE 1st IONIZATION ENERGY tends to ____DECREASE_______

THE ELECTRONEGATIVITY tends to __DECREASE____________

17. The valence electrons are the electrons that are in the __OUTER________energy level. If you are asked to identify how many valence electrons an atom has, all you have to do is count from left to right across the periodic table. Fill in the valence electrons in each box below:



18. The valence electrons will also be written as dots around the atoms. This is called the Lewis dot structure for an atom. Fill in the dots around each atom below:

Li   Be   B   C   N   O   F   Ne

Sometimes an SOL question may give you two different isotopes for an element and ask you to calculate the average atomic mass. Here is an example:

Isotope	Percent abundance
Cl-35	75%
Cl-37	25%

Average atomic mass of Cl = (0.75)(35) + (0.25)(37) = 35.5 amu
(atomic mass units)

19. Now you try it. Calculate the average atomic mass of Cu, based on the data below:

Isotope	Percent abundance
Cu-63	70%
Cu-65	30%

Average atomic mass of Cu = __(0.70)(63) + (0.30)(65) = 64__

20. In Group 1, the most reactive element would be __Cs or Fr__. This can be explained because because metals need to __LOSE__ electrons when they undergo chemical reactions, and so the __LARGER__ the atom, the more reactive it will be.

In Group 17, the most reactive element would be __F__. This can be explained because nonmetals need to __GAIN__ electrons when they undergo chemical reactions, and so the __SMALLER__ the atom, the more reactive it will be.

NOMENCLATURE, CHEMICAL FORMULAS, AND REACTIONS

1. The two main types of bonds in chemistry are __IONIC__ and __COVALENT__
2. An *ionic* bond is normally formed between a __METAL__ and a __NONMETAL__

In an ionic bond, the two elements should have a rather __LARGE__ difference in their electronegativity values. In an ionic bond, electrons are transferred from the __METAL__ to the __NONMETAL__. A classic example of an ionic compound is an alkali metal and a halogen, like NaCl. If an ionic compound is soluble in water, then it will produce aqueous ions in solution. Ionic compounds are considered to be *electrolytes*.

3. Fill in the names and formulas for the following ionic compounds

Chemical Formula	Chemical Name
Na_2S	**SODIUM SULFIDE**
$MgCl_2$	**MAGNESIUM CHLORIDE**
Al_2O_3	**ALUMINUM OXIDE**
Li_3N	**LITHIUM NITRIDE**
K_3P	**POTASSIUM PHOSPHIDE**
CaF_2	calcium fluoride
SrI_2	strontium iodide
CuBr	copper(I) bromide
$CuBr_2$	copper(II) bromide
Fe_2O_3	iron(III) oxide

Remember that we use Roman numerals to indicate the *charge* on the ion when it can form more than one charge.

4. The ionic compounds above are called binary compounds, because they consist of only two elements. Some ionic compounds contain more than two elements. That is because they contain polyatomic ions. The names, formulas, and charges for the following polyatomic ions should be memorized:

ammonium **NH_4^+** carbonate **CO_3^{-2}** nitrate **NO_3^-**

hydroxide **OH^-** sulfate **SO_4^{-2}** phosphate **PO_4^{-3}**

5. Fill in the names and formulas for the following ionic compounds that contain polyatomic ions

Chemical Formula	Chemical Name
$NaNO_3$	**SODIUM NITRATE**
$Fe_2(SO_4)_3$	**IRON(III) SULFATE**
NH_4Cl	**AMMONIUM CHLORIDE**
K_2CO_3	potassium carbonate
$Mg_3(PO_4)_2$	magnesium phosphate
$Ca(OH)_2$	calcium hydroxide

6. A *covalent* bond is normally formed between two **NONMETALS**

In a covalent bond, the two elements should have a rather **SMALL** difference in their electronegativity values. In a covalent bond, electrons are shared between the atoms. A classic example of this is H_2O. If a covalent compound (like sugar, $C_6H_{12}O_6$) is soluble in water, then it will *not* produce any ions. Covalent (molecular) compounds are *nonelectrolytes*.

7. Fill in the names and formulas for the following covalent compounds

Chemical Formula	Chemical Name
CCl_4	carbon tetrachloride
PBr_3	**PHOSPHORUS TRIBROMIDE**
SF_6	**SULFUR HEXAFLUORIDE**
P_2O_5	diphosphorus pentoxide
CS_2	carbon disulfide

Remember that we use prefixes to indicate the *number of atoms* in a covalent compound.

8. The following compounds are classified as *acids*, because they can all *donate H^+*. Write the formulas in the blanks provided. You should know these formulas.

HCl hydrochloric acid **H_2SO_4** sulfuric acid **H_2CO_3** carbonic acid

HNO_3 nitric acid **H_3PO_4** phosphoric acid

You should know that an *acid is a H^+ donor* and it will have a pH that is **LESS THAN 7**

You should know that a base will *accept H^+* and it will have a pH that is **MORE THAN 7**

Examples of acids are listed above. Examples of bases would be anything that contains the hydroxide ion. For example: NaOH, KOH, $Mg(OH)_2$, $Al(OH)_3$, etc.

The *molecular formula* only tells you the number of each kind of atom.

The *structural formula* will also tell you how the atoms are connected to each other.

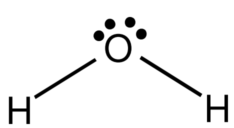
10. The *empirical formula* is the lowest whole number ratio of atoms. For example, the empirical formula of $C_6H_{12}O_6$ is CH_2O. Write the empirical formula for each of the following:

C_6H_{12} **CH_2** $C_{10}H_{20}O_2$ **$C_5H_{10}O$** C_2H_6 **CH_3**

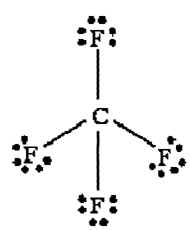
11. Here are some Lewis dot structures for simple molecules. Indicate the geometric shape for each molecule.

Your choices are bent, linear, trigonal planar, pyramidal, tetrahedral.

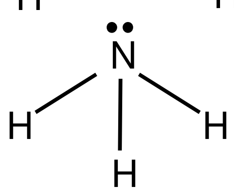
If you learn these 5 examples you should be in very good “shape”.



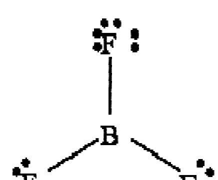
BENT

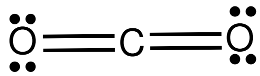


TETRAHEDRAL



PYRAMIDAL



 **LINEAR** **TRIGONAL PLANAR**

12. Balancing equations is a skill that every chemistry student should know how to do. Here are some equations for you to balance.

C_3H_8	+	**5** O_2	→	**3** CO_2	+	**4** H_2O
CH_4	+	**2** Cl_2	→	CH_2Cl_2	+	**2** HCl
$Al(OH)_3$	+	H_3PO_4	→	$AlPO_4$	+	**3** H_2O
2 $FeCl_3$	+	**3** Na_2CO_3	→	$Fe_2(CO_3)_3$	+	**6** $NaCl$
2 Al	+	**3** H_2SO_4	→	$Al_2(SO_4)_3$	+	**3** H_2

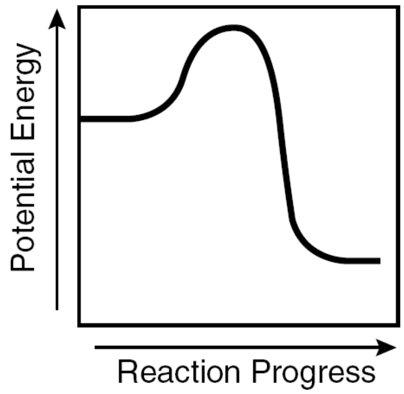
When we learned about chemical reactions, we also learned about that there are categories that describe reaction types. You should be familiar with the following types of reactions:

Reaction type	General Scheme	Specific Example
synthesis	$A + B \rightarrow AB$	$N_2 + 3\,H_2 \rightarrow 2\,NH_3$
decomposition	$AB \rightarrow A + B$	$2\,KClO_3 \rightarrow 2\,KCl + 3\,O_2$
single replacement	$A + BY \rightarrow AY + B$	$Mg + 2\,HCl \rightarrow MgCl_2 + H_2$
double replacement	$AX + BY \rightarrow AY + BX$	$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
neutralization	$HX + MOH \rightarrow H_2O + MX$	$HCl + NaOH \rightarrow H_2O + NaCl$

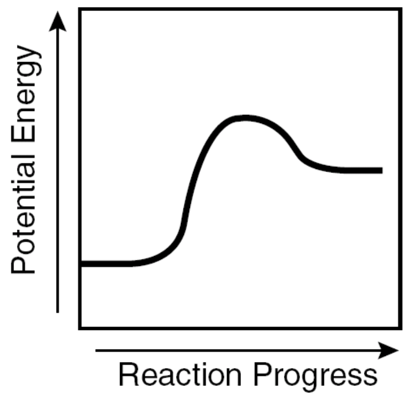
13. Identify the type of each reaction below:
(synthesis, decomposition, single replacement, double replacement, neutralization)

Reaction	Reaction type
$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$	**SINGLE REPLACEMENT**
$HNO_3 + KOH \rightarrow KNO_3 + H_2O$	**NEUTRALIZATION**
$Mg + N_2 \rightarrow Mg_3N_2$	**SYNTHESIS**
$Cl_2 + 2\,NaBr \rightarrow Br_2 + 2\,NaCl$	**SINGLE REPLACEMENT**
$Pb(NO_3)_2 + 2\,KI \rightarrow PbI_2 + 2\,KNO_3$	**DOUBLE REPLACEMENT**
$2\,NH_4NO_3 \rightarrow 2\,N_2 + O_2 + 4\,H_2O$	**DECOMPOSITION**
$Ca(OH)_2 + HBr \rightarrow H_2O + CaBr_2$	**NEUTRALIZATION**
$CaCO_3 \rightarrow CO_2 + CaO$	**DECOMPOSITION**
$K_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2\,KOH$	**DOUBLE REPLACEMENT**

14. Sometimes questions will discuss the energy in a chemical reaction. Here are some things you should know about energy:



This reaction represents an ___EXOTHERMIC___ process. This means that energy is ___RELEASED___ You can think of energy as one of the *products* of the reaction, which means that you would write it on the ___RIGHT___ side of the equation. "ΔH," which represents the change in heat, will be NEGATIVE.

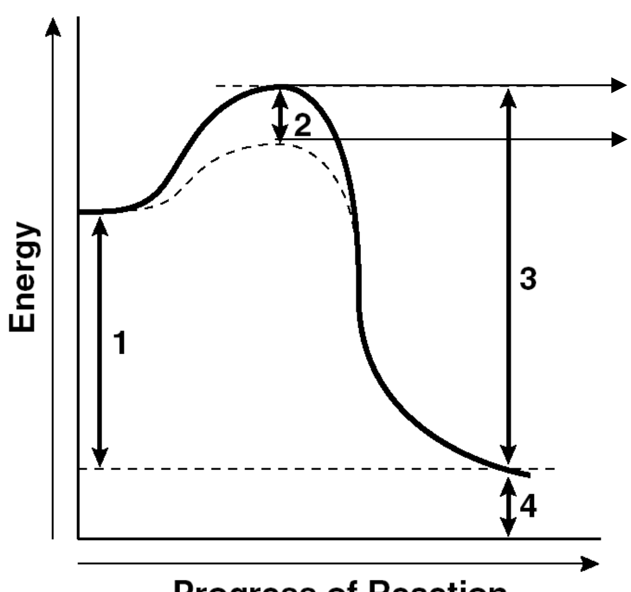


This reaction represents an ___ENDOTHERMIC___ process. This means that energy is ___ABSORBED___. You can think of energy as one of the *reactants*, which means that you would write it on the ___LEFT___ side of the equation. "ΔH," which represents the change in heat, will be POSITIVE.

15. Sometimes you will see a question about how a catalyst speeds up a reaction. Here is an important fact you should know about a catalyst:

A catalyst will ___LOWER___ the activation energy, which makes the reaction go faster.

16. Which path is the catalyzed reaction? Which path is the UNcatalyzed reaction? Label them.



___UNCATALYZED___

___CATALYZED___

17. How do you think the *speed* of the reaction will be affected by temperature? Well, you already know that if you *increase* the temperature, the molecules will move __FASTER__. So if you *increase* the temperature, the molecules will collide with each other more often and the reaction rate will __INCREASE__.

MOLAR RELATIONSHIPS

Here are several important facts about moles that you should know.
You should *memorize* the number 6.02×10^{23} and the number 22.4 and know when to use them.

General Facts	Specific Examples
1 mole = 6.02×10^{23} particles	1 mol of Cu = 6.02×10^{23} atoms of Cu 1 mol of CO_2 = 6.02×10^{23} molecules of CO_2
The mass of 1 mole (in grams) can be calculated by adding up the atomic masses of all the elements in the chemical formula.	1 mol of H_2O = 1.0 + 1.0 + 16.0 = 18.0 g 1 mol of CO_2 = 12.0 + 16.0 + 16.0 = 44.0 g 1 mol of NaCl = 23.0 + 35.5 = 58.5 g
At standard temperature and pressure (STP), 1 mole of gas has a volume of 22.4 L	1 mol of He @ STP = 22.4 L 1 mol of N_2 @ STP = 22.4 L
The coefficients in a balanced chemical equation represent molar ratios.	In the equation $N_2 + 3\ H_2 \rightarrow 2\ NH_3$, this can be summarized by saying that "1 mol of N_2 reacts with 3 mol to produce 2 mol NH_3"

When we perform conversions with moles, we usually set up ratios, or conversion factors, that help us to cancel out units. Remember the following:

- The units will cancel out when they are on opposite sides of the line.
- When a number is above the line we multiply; when a number is below the line we divide.

Here are some examples:

Convert 3.58×10^{24} atoms Fe into moles of Fe

$$3.58 \times 10^{24}\ \cancel{\text{atoms Fe}} \times \frac{1\ \text{mol Fe}}{6.02 \times 10^{23}\ \cancel{\text{atoms Fe}}} = 5.95\ \text{mol Fe}$$

Convert 2.25 moles of KNO_3 into grams of KNO_3 (Note that we need the periodic table to do this.)

K = 39.1 x 1 = 39.1
N = 14.0 x 1 = 14.0
O = 16.0 x 3 = <u>48.0</u>
101.1 g/mol

$$2.25\ \cancel{\text{mol } KNO_3} \times \frac{101.1\ \text{g } KNO_3}{1\ \cancel{\text{mol } KNO_3}} = 227\ \text{g } KNO_3$$

Now it's your turn:

1. Perform the following conversions:

a) Calculate the molar mass of $Ca(NO_3)_2$

Ca = 40.0 x 1 = 40.0
N = 14.0 x 2 = 28.0
O = 16.0 x 6 = <u>96.0</u>
164.0 grams per mole

b) How many *grams* of oxygen are present in 2 moles of $CaCO_3$?

$$\textbf{2 moles of } CaCO_3 \times \frac{3 \text{ mol O}}{1 \text{ mol } CaCO_3} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}} = 96.0 \text{ g O}$$

c) How many *moles* are present in a 100.0-g sample of C_2H_6O?

$$100.0 \text{ g } C_2H_6O \times \frac{1 \text{ mol } C_2H_6O}{46.0 \text{ g } C_2H_6O} = 2.17 \text{ mol } C_2H_6O$$

d) What is the *mass* of 9.25×10^{22} molecules of water? (two steps).

$$9.25 \times 10^{22} \text{ molecules } H_2O \times \frac{1 \text{ mol } H_2O}{6.02 \times 10^{23} \text{ molecules } H_2O} \times \frac{18.0 \text{ g } H_2O}{1 \text{ mol } H_2O} = 2.77 \text{ g } H_2O$$

Another type of molar conversion you will be asked to do is related to a balanced chemical equation. We will use the coefficients to set up molar ratios, or conversion factors. Again we will try to cancel out units. Here is an example:

When magnesium metal is burned, it produces magnesium oxide (MgO). How many moles of oxygen gas are needed to burn 10 moles of Mg?

In this problem we are *not* given a balanced chemical equation, so we have to write one first:

here is the equation: $Mg + O_2 \rightarrow MgO$

and now it is balanced: $2\,Mg + O_2 \rightarrow 2\,MgO$

Notice that there are 2 moles of Mg for every 1 mole of O_2. That is the molar ratio you need.

$$10 \cancel{\text{ mol Mg}} \times \frac{1 \text{ mol } O_2}{2 \cancel{\text{ mol Mg}}} = 5 \text{ mol } O_2$$

Now it's your turn:

2. Perform the following conversions:

a) Given the following equation: $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$
If 5.2 moles of ethane (C_2H_6) is burned, how many moles of O_2 are required?

$$\mathbf{5.2\ mol\ C_2H_6 \times \frac{7\ mol\ O_2}{2\ mol\ C_2H_6} = 18.2\ mol\ O_2}$$

b) Given the following equation: $2\ Al + 6\ HCl \rightarrow 2\ AlCl_3 + 3\ H_2$
If 3.4 moles of aluminum reacts with excess hydrochloric acid, how many moles of H_2 will be produced?

$$\mathbf{3.4\ mol\ Al \times \frac{3\ mol\ H_2}{2\ mol\ Al} = 5.1\ mol\ H_2}$$

Sometimes you are asked to convert *grams of one chemical into grams of another chemical*. With this type of molar conversion you will need to do *three steps*.
Again we will try to cancel out units. Here is an example from the 2005 SOL test:

$$\boxed{\mathbf{2\ KOH + H_2SO_4 \rightarrow 2\ H_2O + K_2SO_4}}$$

What mass of potassium hydroxide is required to react completely with 2.70 g of sulfuric acid to produce potassium sulfate and water?

In this problem you need to go from *grams of H_2SO_4 into grams of KOH*

If you take it one step at a time, and remember to set up the units so they will cancel out, then this is not a difficult problem:

This is the basic set-up, with the units in place. Notice how everything cancels out except for the grams of KOH at the end of the problem.

$$2.70\ g\ H_2SO_4 \times \frac{mol\ H_2SO_4}{g\ H_2SO_4} \times \frac{mol\ KOH}{mol\ H_2SO_4} \times \frac{g\ KOH}{mol\ KOH} = \qquad g\ KOH$$

The 1st step requires the periodic table. When we add up all the atomic masses for H_2SO_4, we get (2)(1.0) + (32.0) + (4)(16.0) = 98.0 g/mol

The 2nd step requires the coefficients. We see that 2 moles of KOH react with 1 mole of H_2SO_4.

The 3rd step requires the periodic table again. When we add up all the atomic masses for KOH, we get (39.1) + (16.0) + (1.0) = 56.1 g/mol

Now put all the numbers in place:

$$2.70\ g\ H_2SO_4 \times \frac{1\ mol\ H_2SO_4}{98.0\ g\ H_2SO_4} \times \frac{2\ mol\ KOH}{1\ mol\ H_2SO_4} \times \frac{56.1\ g\ KOH}{1\ mol\ KOH} = 3.09\ g\ KOH$$

Remember that if a number is above the line you multiply and if it is below the line you divide.

Now it's your turn:

3. Perform the following conversion:

   Given the following equation: $Pb(NO_3)_2 + 2\,KI \rightarrow PbI_2 + 2\,KNO_3$
   If 5.00 grams of potassium iodide reacts according to the equation above, how many grams of lead iodide will be produced?

$$5.00\text{ g KI} \times \frac{1\text{ mol KI}}{166.0\text{ g KI}} \times \frac{1\text{ mol }PbI_2}{2\text{ mol KI}} \times \frac{461.0\text{ g }PbI_2}{1\text{ mol }PbI_2} = 6.94\text{ g }PbI_2$$

Another type of problem that you will need to know involves moles of gas at standard temperature and pressure (STP). Conditions of STP are pressure = 1 atm and temp. = $0^{o}C$
The equation we use for gases is called the ideal gas law: PV = nRT
P = pressure, V = volume, n = moles, R = a gas constant, and T = temperature.
When you solve for the volume of 1 mole of any gas at STP, this is what you get:

$$V = \frac{nRT}{P} = \frac{(1\text{ mol})(8.31\text{ kPa L mol}^{-1}\text{ K}^{-1})(273\text{ K})}{(101.3\text{ kPa})} = 22.4\text{ L}$$

Because this number is used so often, you should just memorize that

1 mole of any gas at STP has a volume of 22.4 L

There are a number of ways in which you can use this information. Try the following examples:

4. a) What is the density of CH_4 gas at STP?

   1 mol CH_4 = 16.0 g and also has a volume of 22.4 L, so density = $\frac{16.0\text{ g}}{22.4\text{ L}} = 0.714\text{ g/L}$

   b) Which sample of gas has the largest volume at STP?

   10.0 g He   10.0 g Ne   10.0 g Ar   10.0 g Kr

   The sample with the largest amount of moles should have the largest volume.
   To convert from grams to moles we need to divide 10.0 g by the molar mass for each gas. That means that Helium is the sample with the largest # of moles.

   c) What is the volume of 3.01×10^{23} atoms of He gas at STP?

$$3.01 \times 10^{23}\text{ atoms He} \times \frac{1\text{ mol He}}{6.02 \times 10^{23}\text{ atoms He}} \times \frac{22.4\text{ L}}{1\text{ mol He}} = 11.2\text{ L He}$$

$$N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)$$

Suppose that you have a balanced chemical equation like the one above, and that all of the chemicals are gases. You already know that the coefficients represent *molar ratios*. But if all of the chemicals are gases, then the coefficients *also represent volume ratios!*
In other words, the equation above can be thought of as the following:
"1 liter of N_2 reacts with 3 liters of H_2 to produce 2 liters of NH_3."

Suppose the question asks you something like the following:

How many liters of hydrogen gas are needs to react completely with 2.00 L of nitrogen gas?

All you have to do is use the coefficients as the ratio between liters of H_2 and liters of N_2:

$$2.00 \text{ L } N_2 \times \frac{3 \text{ L } H_2}{1 \text{ L } N_2} = 6.00 \text{ L } H_2$$

It's really simple. Try the following examples:

5. a) Given the following equation: $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(g)$
To produce 12 liters of water, how many liters of oxygen gas are needed?

$$\mathbf{12 \text{ L } H_2O \times \frac{7 \text{ L } O_2}{6 \text{ L } H_2O} = 14 \text{ L } O_2}$$

b) Given the following equation: $2\ H_2S(g) + 3\ O_2(g) \rightarrow 2\ H_2O + 2\ SO_2(g)$
If 4.0 liters of oxygen gas reacts according to the above reaction, how many liters of H_2S will be required?

$$\mathbf{4.0 \text{ L } O_2 \times \frac{2 \text{ L } H_2S}{3 \text{ L } O_2} = 2.7 \text{ L } H_2S}$$

The last topic in molar relationships deals with molarity (M), which is defined as follows:

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

This equation can be rearranged: (M) x (liters of solution) = moles

Remember the following:

If you are given *grams* of solute, you can convert it into *moles* using the periodic table.
Of course you can also go from moles to grams, too.

If you are given a volume in *mL*, you can convert it into *liters* by dividing by 1000. For example, 500 mL = 0.500 L. Of course you can also go from liters to mL by multiplying by 1000.

Here are some example problems that deal with molarity.

6. a) How many grams of KCl are required to prepare 500 mL of a 0.125 M solution?

The molar mass of KCl is equal to

$$\mathbf{0.500 \text{ L} \times \frac{0.125 \text{ mol KCl}}{} \times \frac{74.6 \text{ g KCl}}{} = 4.66 \text{ g KCl}}$$

39.1 + 35.5 = 74.6 g/mol **1 L** **1 mol KCl**

b) What is the molarity of a solution that is prepared by dissolving 75.0 g of $C_6H_{12}O_6$ in enough water to prepare 500.0 mL of solution?

C = 12.0 x 6 = 72.0
H = 1.0 x 12 = 12.0
O = 16.0 x 6 = 96.0
180.0 g/mol

$$75.0 \text{ g} \times \frac{1 \text{ mol } C_6H_{12}O_6}{180.0 \text{ g } C_6H_{12}O_6} = 0.417 \text{ mol } C_6H_{12}O_6$$

Molarity = (moles / L) = (0.417 mol / 0.500 L) = 0.833 M

c) How many milliliters of 2.50 M NaCl are needed to provide 0.150 mol NaCl?

$$0.150 \text{ mol NaCl} \times \frac{1 \text{ L}}{2.50 \text{ mol NaCl}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 60 \text{ mL}$$

Sometimes a solution is prepared by diluting (adding water) to a concentrated solution. If you have to do a problem that involves dilution, here is how you do it:

$M_1V_1 = M_2V_2$ where M_1 is the initial molarity of the concentrated solution
M_2 is the final molarity of the diluted solution
V_1 is the initial volume of the concentrated solution
V_2 us the final volume of the diluted solution

Example: *A 15 mL sample of 4.0 M NaOH was diluted to a volume of 250 mL. What is the new concentration of the solution?*

(4.0 M)(15 mL) = (M_2)(250 mL)

$$M_2 = \frac{(4)(15)}{250} = 0.24 \text{ M}$$

7. a) If 50.0 mL of a 3.00 M solution is diluted to a volume of 500 mL, what is the final concentration?

(3.00 M)(50.0 mL) = (M_2)(500 mL)

$$M_2 = \frac{(3)(50)}{500} = 0.30 \text{ M}$$

Note that the volume increased **ten-fold and the molarity** decreased **by a factor of ten.**

b) 750 mL of 0.50 M HCl is required for a lab experiment. How many milliliters of 6.00 M HCl should be used to prepare this solution?

(6.00 M)(V_1) = (0.50 M)(750 mL)

$$V_1 = \frac{(0.5)(750)}{6} = 62.5 \text{ mL}$$

PHASES OF MATTER AND KINETIC MOLECULAR THEORY

Here are some important things to know about the kinetic molecular theory:

Gas particles are in constant, rapid, random motion, and they are very far part from each other. When you increase the temp., gas particles travel faster because they have more kinetic energy.

Here are some gas laws you should know:

Charles' Law: As temp. goes up, volume goes up (and vice versa)
Boyle's Law: As pressure goes up, volume goes down (and vice versa)

If you ever see a problem involving a gas collected "by water displacement" or "over water," you will always *subtract the water pressure from the total pressure to get the pressure of the dry gas.*

1. For example: A sample of oxygen gas is collected over water at 98.67 kPa. If the partial pressure of the water is 2.67 kPa, the partial pressure of the oxygen is **96.00 kPa**

In general, the total pressure of a gas *mixture* is equal to the sum of the partial pressures of each individual gas.

2. If you have to do any calculations with gases that involve temperature, you should always convert the temperature from ^{o}C to K by **ADDING 273 TO IT**

Here is an example: A sample of gas occupies a volume of 5.00 L at 25°C. This gas was heated at constant pressure and the volume increased to 6.00 L. What is the new temperature of the gas?

Charles Law: $\frac{T_1}{V_1} = \frac{T_2}{V_2}$ $\qquad \frac{(298\ K)}{(5.00\ L)} = \frac{(T_2)}{(6.00\ L)}$ $\qquad T_2 = \frac{(6.00)(298)}{(5.00)} = 358\ K - 273 = 85^{o}C$

3. A sample of gas occupies a volume of 10.0 liters at 10°C. What would be the volume of this gas at 50°C if the pressure remains constant?

Charles Law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $\qquad \frac{(10.0\ L)}{(283\ K)} = \frac{(V_2)}{(323\ K)}$ $\qquad V_2 = \frac{(10.0)(323)}{(283)} = 11.4\ L$

If you are given a problem that involves the ideal gas law, you will need to remember

$$PV = nRT$$

Here is an example problem:

$$\boxed{R = 8.31\ \frac{kPa \bullet dm^3}{moles \bullet K}}$$

A gas cylinder is filled with 4.00 moles of oxygen gas at 300.0 K. The piston is compressed to yield a pressure of 400.0 kPa. What is the volume inside the cylinder?

A 3.19 dm^3
B 6.25 dm^3
C 24.9 dm^3
D 31.5 dm^3

To answer this question, you need to solve for V:

$$V = \frac{nRT}{P} = \frac{(4.00 \text{ moles})(8.31 \text{ kPa dm}^3 \text{ mol}^{-1} \text{ K}^{-1})(300.0 \text{ K})}{(400.0 \text{ kPa})} = 24.9 \text{ dm}^3$$

(Notice that all units cancel out except dm^3)

$$\boxed{R = 8.31 \frac{\text{kPa L}}{\text{mol K}}}$$

4. A sample of oxygen gas occupies a volume of 15.0 liters at a pressure of 250 kPa and a temperature of $50^{o}C$. How many *moles* of oxygen are present in this gas sample?

$$\mathbf{n = \frac{PV}{RT} = \frac{(250 \text{ kPa})(15.0 \text{ L})}{(8.31 \text{ kPa dm}^3 \text{ mol}^{-1} \text{ K}^{-1})(323 \text{ K})} = 1.40 \text{ moles}}$$

5. There are other things you should know about phases of matter:
Fill in the name of the phase changes below:

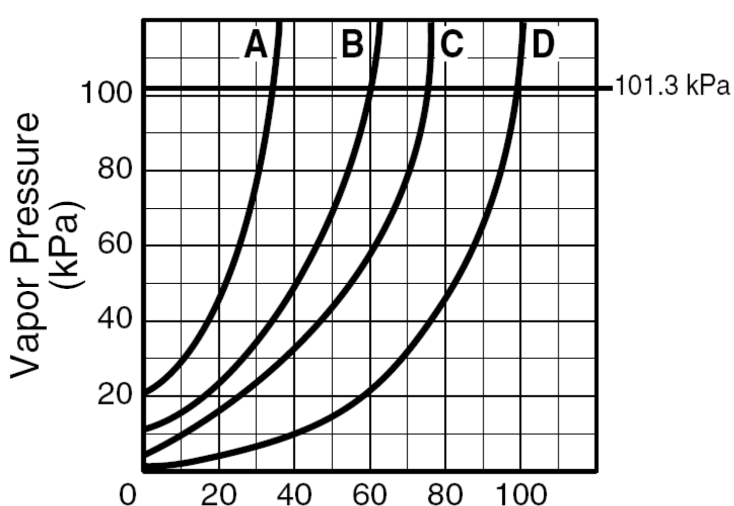
These three phase changes are all ENDOTHERMIC:					
Solid → Liquid	**MELTING**	**Liquid → Gas**	**EVAPORATION**	**Solid → Gas**	**SUBLIMATION**

These three phase changes are all EXOTHERMIC:					
Gas → Liquid	**CONDENSATION**	**Liquid → Solid**	**FREEZING**	**Gas → Solid**	**DEPOSITION**

Another word for melting is FUSION. Another word for evaporation is VAPORIZATION.

If you see a diagram with a sealed liquid in a jar or flask, you should know that there is an equilibrium happening in there. The rate of evaporation is equal to the rate of condensation.

Vapor pressure is defined as the pressure exerted by the gas above a liquid.
Here is an example of some vapor pressure curves:



6. From this graph we can get certain information.

   a) The normal boiling point of liquid A is __**34-35°C**______

   b) If the external pressure is reduced to 60 kPa, then Liquid C would boil at **60-61°C**___

   c) The liquid with the strongest intermolecular forces is most likely **LIQUID D**__

Liquid	Boiling Point (°C)
ether	35
ethyl alcohol	78
water	100
glycerine	290

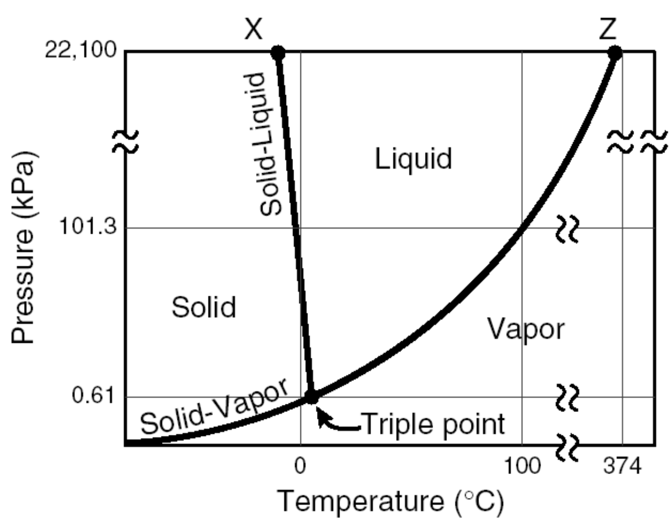
7. Which of the liquids in the table above would have the *highest vapor pressure* at room temperature? Explain

ETHER HAS THE LOWEST BOILING POINT. THEREFORE IT SHOULD HAVE THE HIGHEST VAPOR PRESSURE AT ROOM TEMPERATURE. A LIQUID WITH A HIGH VAPOR PRESSURE SHOULD EVAPORATE EASILY, AND ETHER PROBABLY HAS THE WEAKEST IATTRACTIVE FORCES.

8. If you want to get water to boil BELOW 100°C, you can ____**DECREASE**___ the air pressure.

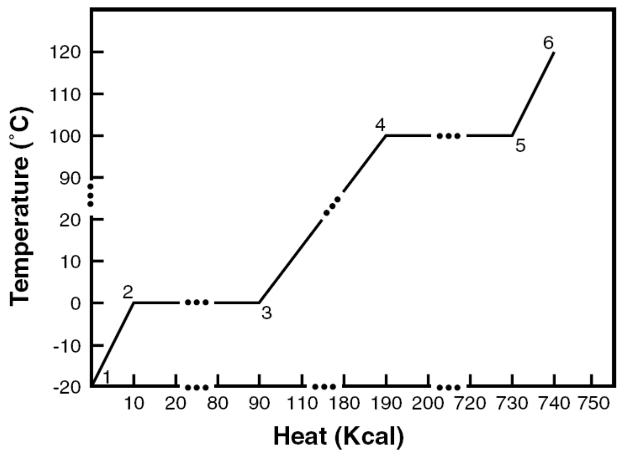
   If you want to get water to boil ABOVE 100°C, you can ____**INCREASE**___ the air pressure.

9. If you add salt to water, this will **LOWER**___ the freezing point and **RAISE**_ the boiling point.



10. The diagram above is called a phase diagram. All along the boundary between two phases there is an equilibrium between those phases. What can we say about the triple point?

AT THE TRIPLE POINT, ALLTHREE PHASES OF MATTER (SOLID, LIQUID, AND GAS) ARE IN EQUILIBRIUM WITH EACH OTHER.



11. The diagram above is called a heating curve. Match the descriptions of what is happening with the various line segments

C	Between 1 and 2	A. ice is melting
A	Between 2 and 3	B. liquid water is evaporating
D	Between 3 and 4	C. ice is being heated
B	Between 4 and 5	D. liquid is being heated
E	Between 5 and 6	E. gas is being heated

Sometimes you will be asked to calculate how much heat is needed to raise the temperature of water. Here is an example:

How many calories of heat are needed to raise the temperature of 50.0 g of water from 20.0°C to 80.0°C?

You should know that it takes ONE CALORIE to raise the temperature of ONE GRAM of water by ONE DEGREE CELSIUS.

So all you have to do is use the following equation:

$(\text{MASS}) \times \left(\frac{1 \text{ calorie}}{\text{g } ^\circ\text{C}}\right) \times (\Delta T)$ where ΔT is the change in temperature.

(50.0 g) x (1 cal / g °C) x (60°C) = 3000 calories

12. How many calories are needed to raise the temperature of 75.0 g H_2O from 30.0°C to 70.0°C?

(75.0 g) x (1 cal / g °C) x (40°C) = 3000 calories (OOH, THAT WAS JUST A COINCIDENCE!)

Sometimes you will be asked to calculate how much heat is needed to melt a substance. They will give you the heat of fusion. Here is an example:

13. The heat of fusion for water is 6.12 kJ per mole. How many kJ of heat is required to melt 100.0 grams of ice at 0^oC?

$$100.0 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O} \times \frac{6.12 \text{ kJ}}{1 \text{ mol } H_2O} = 34 \text{ kJ}$$

14. If you see a question that mentions that water has a high boiling point or a high heat capacity, then the explanation will be that water has very strong intermolecular forces, known as

___HYDROGEN BONDING________________

15. If you see any questions that deal with polarity and mixing two liquids together, you should know that two liquids will mix well together if they are **___BOTH POLAR OR BOTH NONPOLAR**

16. You might be asked to predict if the attractive forces are strong or weak. You should know that if a substance has a high melting or boiling point, then it will have **___STRONG______** attractive forces.

END OF SOL REVIEW PACKET. GOOD LUCK ON THE SOL TEST!