Name __________________   **Chemistry**   **___/___/___**

Honors Chemistry Final Exam Review 2021

Chapter 1 – Matter

Important Formulas:

$$\text{Density(solid \& liquid)} = \frac{\text{Mass(g)}}{\text{Volume(mL)}}$$

$$\%\ \text{Error} = \frac{|\text{theoretical} - \text{experimental}|}{\text{theoretical}} \times 100 \qquad \%\ \text{Yield} = \frac{\text{experimental}}{\text{theoretical}} \times 100$$

Lab Info:

- All measurements must be made using the same balance.
- All measurements should be made numerous times to ensure results.
- In all measurements you should estimate a digit beyond the precision of the instrument used. (For example if the ruler is divided into tenths of a centimeter, estimate to the hundredths place.)
- Do as you otta, add Acid to Water!!! (Italian-American) or Do what is proper, add acid to water. (British)
- If chemicals ever come in contact with skin or eyes, rinse with WATER!!! Do not neutralize. Bases are worse for skin.
- Point test tubes AWAY from others when heating.
- Do not use a stopper when heating a test tube. It will pop out and potentially hit someone.
- Waft chemicals towards your nose. Remember ammoniun carbonate (#2)?
- All measurements of a liquid VOLUME must be done in a GRADUATED CYLINDER!!!
- Read liquid volume at the bottom of the meniscus.
- Beakers and flasks are for storage and heating of liquids.
- A "control" is part of an experiment where none of the tested variables are used. It is used as a reference to measure the effect of the variable on the experiment.
- Know how to identify the following lab equipment: beaker, flask, test tube, test tube rack, ring stand, wire gauze, clay triangle, crucible, watch glass, evaporating dish.

General Info:

- **Accuracy**: how close a number is to the **actual** scientific result.
- **Precision**: having **repeated results** of the same value.
- Scientific notation: One number before the decimal. (1.23×10^3 YES) (12.3×10^2 NO!)
- Move your decimal left and your power of ten goes up. Move your decimal to the right and your power of ten goes down. Think Lift(Left) Up, Write(Right) Down
- Metric units to know: Kilo (1×10^3), Base Units(no prefixes) & Milli (1×10^{-3}). Move Up (milli → Base unit, etc.) divide by 1000. Move down (kilo → base unit, etc.) multiply by 1000.
- The mean of a set of numbers is the average.
- Significant figures. All non-zero numbers are significant. Zeros at the beginning of a number are not significant. Zeros at the end of a number are significant if there is a decimal.
- Calculations using significant figures: A SAD – MODE: **a**dd or **s**ubtract **a**fter **d**ecimal – **m**ultiply **o**r **d**ivide **e**verything
- Physical Properties - properties that can be observed or measured without changing the composition of the matter.
- Chemical Properties - properties that can be observed only when substances change their composition.

Ways to Separate a Mixture:

- Filtration – separating by particle size, large particles can not travel through a filter.
- Chromatography – separation by color
- Distillation – separation by boiling point (collect the substance with the lowest boiling point)
- Evaporation – separation by boiling point (left with substance with the higher boiling point)
- Decanting – separation where the top liquid is poured off leaving the solid behind
- Sublimation – separation where one substance changes from solid directly to gas without becoming a liquid

Chapter 2 – Atomic Theory

Subatomic Particles:

- **Proton (+) found in nucleus**; the nucleus was discovered by **Rutherford** in the **Gold Foil Experiment**
- **Neutron (neutral)**; found in nucleus; heaviest subatomic particle
- **Electron (-)** found **outside the nucleus**; lightest subatomic particle; 1800 electrons = 1 proton = 1 neutron; discovered by **J.J. Thomson** when he used a magnet on a cathode ray tube

Scientists:

- Dalton – Father of the Atomic Theory; developed first atomic theory; Dalton's Law of Partial Pressure
- Thomson – discovered electron while using a magnet with a cathode ray tube; Plum Pudding Model
- Rutherford – discovered nucleus is a small dense positively charged core in the atom, atoms are mostly empty space; Gold Foil Experiment
- Bohr – put electrons into energy levels, like planets around the sun; very handsome man
- Planck – all energy is quantized
- Millikan – determined the charge of the electron
- Schrödinger – quantum mechanics model a.k.a. electron cloud model; I like the ö dots
- Mendeleev – Father of the Periodic Table; he was quicker than Meyer
- Heisenberg – the uncertainty principle, you can not determine both the location and velocity of an electron without affecting its path or speed.
- Aufbau Principle - electrons enter orbitals of lowest energy first. (ex. $2s^2$ is filled before $2p^6$)
- Pauli Exclusion Principle - an atomic orbital can describe at most two electrons; those two electrons circle the nucleus in opposite directions
- Hund's Rule - when electrons occupy orbitals of the same energy(for example 2p), electrons will enter empty orbitals first(one in px, one in py and one in pz before you put a second electron in px)

General Info:

- **Atoms of the same element have the same number of protons**. But atoms of the same element can vary in the number of electrons(ions) and neutrons(isotopes) they have.
- **Isotope** – an atom of an element that has the same number of protons but a different number of neutrons and as a result a different mass number. Example: chlorine-35 and chlorine-37 (35 & 37 represent the mass number. The mass number = # p + # n. So chlorine -35 has17 protons and 18 neutrons; chlorine-37 has 17 protons and 20 neutrons.)
- Atomic number = #protons
- Mass number = #protons + #neutrons
- Atomic mass (as it is seen on the periodic table) of an element is the weighted average of that element's naturally occurring isotopes. Example: Chlorine-35 has an abundance of 75%. Chlorine-37 has an abundance of 25%. To find the atomic mass of chlorine you would make the following calculation:
  (35 x 0.75) + (37 x 0.25) = 35.5 If you get a question like this do the math for each and see which one is closest to the actual atomic weight.
- In a neutral atom #protons = # electrons. If an atom gains or loses an electron it is called an ion. Cations have positive(+) charges because they lost electrons. Anions have negative (-) charges because they gain electrons. (Remember the electron is negative.)

Cl	Cl^{1-}	Fe	Fe^{2+}
#p+ = 17	#p+ = 17	#p+ = 26	#p+ = 26
#e- = 17	#e- = 18	#e- = 26	#e- = 24
anion (gains electron (-) charge)		cation (loses electron (+) charge)	

- **Electron configuration:**
  - s: 2 electrons max
  - p: 6 electrons max
  - d: 10 electrons max
  - f: 14 electrons max
- If the test asks which element would have an electron configuration like: ns^2np^3, The n refers to any integer, not important. Look at how it ends…p^3. Go to the p section, and starting on the left side of the p section, count over three columns. All elements in that column (group) have the same ending to their electron configuration.

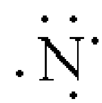
- Nuclear chemistry is the study of the changes in the composition of the nucleus of the atom. Radioisotopes (radioactive isotopes) are used in medical diagnosis and dating ancient materials.
- Alpha particles – helium nuclei , can be stopped with a piece of paper
- Beta particles – electrons, can be stopped by metal foil
- Gamma particles – high energy electromagnetic radiation, lead and concrete are needed to stop gamma particles
- Fusion – combining of two nuclei to produce one heavier nucleus, large amount of energy is released
- Fission – splitting of a nucleus into smaller fragments, lots of energy is released.
- Half-life – the amount of a radioactive substance that remains after a certain period of time. For example a 100 gram substance has a half life of 3 days. So, in three days I would have 50 grams. Three days later (day 6) I would have 25 grams. On day 9 I would have 12.5 grams and on day 12 I would have 6.25 grams.

- COLOR THIS TABLE FOR s,p,d & f --------------------

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Uuu	112 Uub		114 Uuq				
		58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu		
		90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr		

Chapter 3 – Periodic Table

- Shielding effect increases with radius, inner electrons block nucleus from outer electrons.
- Elements on the periodic table are arranged by increasing atomic number. As you move left to right across the periodic table elements gain protons in the nucleus and electrons outside the nucleus.
- Periodicity is regularly repeating patterns or trends in the chemical and physical properties of the elements arranged in the periodic table.
- Vertical columns called groups have **similar chemical properties** because of their similar valence electron configurations.
- Horizontal rows called periods have properties based on an increasing number of electrons in the outer orbitals.
- Periodic Law states that when elements are arranged in order of increasing atomic number, their physical and chemical properties show a periodic pattern.
- Diatomic elements: H_2 N_2 O_2 F_2 Cl_2 Br_2 I_2
- **Valence electrons** are electrons in the outer energy level of an atom. The valence electrons of an element can be determined by the group number it is in. For example nitrogen(N) is in group 15 (VA). Elements in this group have 5 valence electrons. Valence electrons are sometimes written as dots in a Lewis dot diagram.



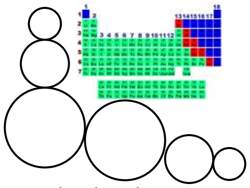
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Uuu	112 Uub		114 Uuq				
		58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu		
		90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr		

metal (green)
metalloid (red)
non-metal (blue)

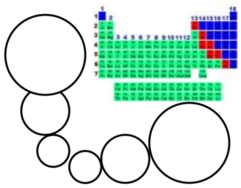
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Uuu	112 Uub		114 Uuq				
		58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu		
		90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr		

alkali metals(red)
alkaline earth (orange)
transition elements (green)
lanthanides (yellow)
actinides (purple)
chalcogens (brown)
halogens (blue)
noble gases (grey)

Remember the snowman for **atomic radius** (half the distance between the nuclei of two atoms located next to each other).

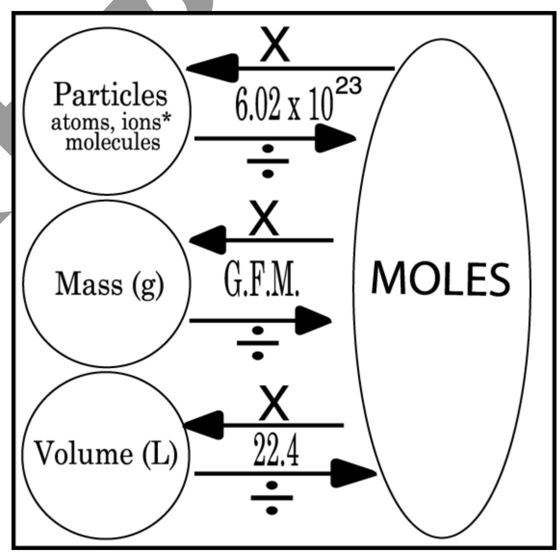


For **ionization energy** (amount of energy needed to remove an electron from an atom in the ground state and gas phase) and **electronegativity** (the tendency to attract electrons from another element when it is bonded to that element), **the pattern is the opposite**.



Chapter 4 – Ionic Compounds

- Ionic compounds are composed of a cation and an anion. The cation is usually a metal and the anion is always a non metal. **Cross charges** when writing a formula.
- If a substance has more than one charge a Roman numeral is used to indicate the oxidation number of the cation. Almost all transition elements need a Roman numeral.
- **To indicate the number of atoms of each element present in a compound, scientists use subscripts.** (i.e. Fe_2O_3 is iron(III) oxide; FeO is iron (II) oxide (reduced).
- Polyatomic Ions to Memorize**: $(CO_3)^{2-}$ carbonate, $(SO_4)^{2-}$ sulfate, $(NO_3)^{1-}$ nitrate, $(PO_4)^{3-}$ phosphate, $(OH)^{1-}$ hydroxide, $(NH_4)^{+}$ ammonium**



Chapter 5 – Chemical Quantities

Mole Conversions:

- **YOU MUST MEMORIZE THIS CHART→**
- Avogadro's number – 6.02×10^{23} particles
- 1 mole of any gas at STP = 22.4 liters

Percent Composition: Calculate the percent composition of the elements in ammonium nitrate.

Description of Action	Action
1. Write the formula for the given compound.	**1.** NH_4NO_3
2. Record the amount of each element in the compound. (Note: We have 2 total nitrogen, so we record them together.) Multiply the amount of each element by its atomic weight (measured to the tenths place). Add the results to find the gram formula mass of the compound.	**2.** N: 2 x 14.0 = **28.0** H: 4 x 1.0 = **4.0** O: 3 x 16.0 = **48.0** **80.0 g/mol**
3. Divide the total mass of each element by the gram formula mass. For these calculations your answer should have 3 places after the decimal (round if necessary).	**3.** N: 28.0 ÷ 80.0 = **0.350** H: 4.0 ÷ 80.0 = **0.050** O: 48.0 ÷ 80.0 = **0.600**
4. Multiply each result by 100. Add the % symbol to your new result. (If you were to add up your percentages the must equal 100.)	**4.** N: 0.350 x 100 = **35.0%** H: 0.050 x 100 = **5.0%** O: 0.600 x 100 = **60.0%** **100.0%**

Empirical Formula: A compound is 36.5% sodium, 25.4% sulfur and 38.1% oxygen. Determine its empirical formula.

Description of Action	Action
1. Divide each element's percent composition by its atomic weight. Remember to use significant figures.	Na: 36.5 ÷ 23.0 = **1.59** S: 25.4 ÷ 32.1 = **0.791** O: 38.1 ÷ 16.0 = **2.38**
2. Divide each result by the **smallest result**. Remember to use significant figures.	Na: 1.59 ÷ 0.791 = **2.01** S: 0.791 ÷ 0.791 = **1.00** O: 2.38 ÷ 0.791 = **3.01**
3. Write the formula with the each element's result as its subscript.	**Na_2SO_3**

- Empirical formulas are reduced to their simplest form.

Molecular Formula

A compound has an empirical formula CH_2O, determine its molecular formula if its molecular mass is 180 g/mol.

Description of Action	Action
1. Determine the gram formula mass of the empirical formula.	**1.** C: 1 x 12.0 = 12.0 H: 2 x 1.0 = 2.0 O: 1 x 16.0 = <u>16.0</u> **30.0 g/mol**
2. Divide the molecular formula mass by the empirical formula mass.	**2.** 180 ÷ 30.0 = 6
3. Multiply each subscript in your empirical formula by your result.	**3.** **$C_6H_{12}O_6$**

Chapter 6 – Chemical Reactions

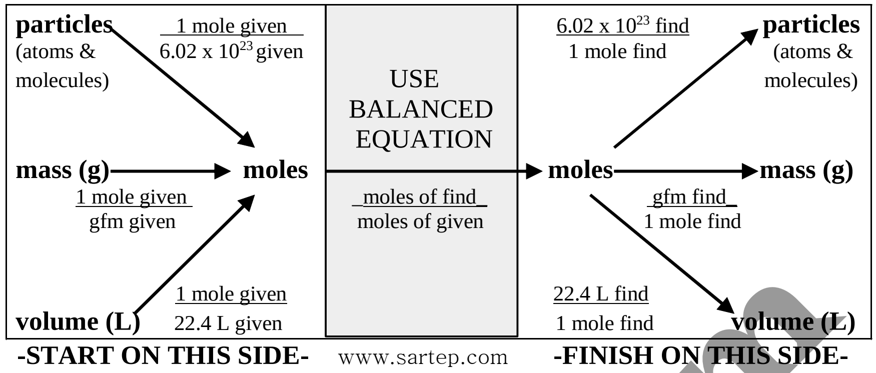
- You will be expected to be able to balance chemical equations. Be sure to try all options!
- Be able to identify the following types of reactions:
  - Synthesis/Combination: A + B → AB — $2Na + Cl_2 \rightarrow 2NaCl$
  - Decomposition: AB → A + B — $MgBr_2 \rightarrow Mg + Br_2$
  - Single Replacement: A + BC → AC + B — $Li + NaCl \rightarrow LiCl + Na$
  - Double Replacement: AB + CD → AD + CB — $KBr + NaCl \rightarrow KCl + NaBr$
  - Combustion: $C_xH_y + O_2 \rightarrow CO_2 + H_2O$ — $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
  - Neutralization: Acid + Base → Salt + Water — $HCl + NaOH \rightarrow NaCl + H_2O$
- Balanced equations show the law of conservation of matter. The mass of the reactants must equal the mass of the products.
- Reactions can be exothermic (release heat) or endothermic (absorb heat)
- Reactions can occur in two directions simultaneously (reversible reactions). They have $\rightleftharpoons$ as a yields sign.
- **Catalysts reduce the amount of activation energy** needed to start a chemical reaction **and speed up chemical reactions** but are not used up in the chemical reaction.

Chapter 7 – Covalent Compounds

- Formed between two non-metals.
- Electrons are **shared** between atoms.
- Each atom needs 8 electrons, except hydrogen, which only needs 2.
- Named using prefixes (mono-, di-, tri-, tetra-, penta-, hexa, hepta-, octa-, nona-, deca-) based on the number of atoms in compound.
- H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2 are diatomic molecules.
- You should know how to write structural formulas. (Commonly used molecules: CH_4, NH_3, H_2O, N_2, O_2, F_2)
- Be familiar with molecular shapes, tetrahedral (ex. CH_4), pyramidal (ex. NH_3), bent (ex. H_2O), linear (N_2, triple bond; O_2, double bond; F_2 single bond), and trigonal planar (ex: BF_3).
- Polar molecules result when a molecule behaves as if one end were positive and the other end negative.
- Non-polar molecules share electrons equally.
- Water is polar, oils are non-polar.
- Forces of attraction between molecules determine the physical changes of state. The stronger the intermolecular forces, the higher the melting and boiling point.
- Intermolecular forces include: hydrogen bonding, dipole interactions and dispersion (London) forces.
- Water's unique properties are due to hydrogen bonding.

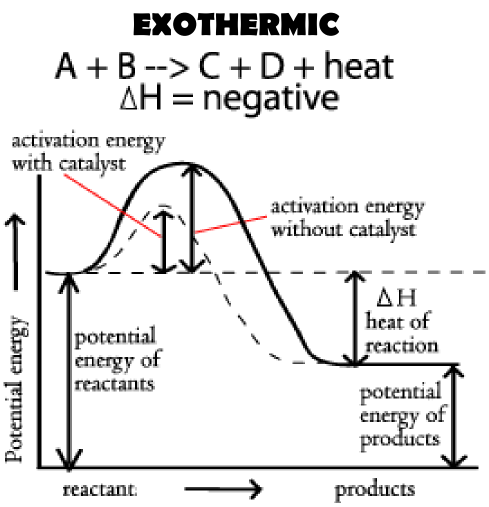
Chapter 8 – Stoichiometry

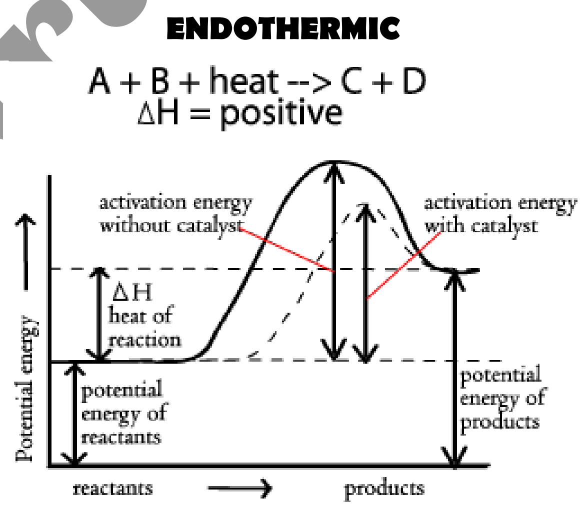
- Expect to see the following: Mole to Mole, Liter to Liter, Molecule to Molecule and Mass to Mass Problems. You must memorize the chart to the right.



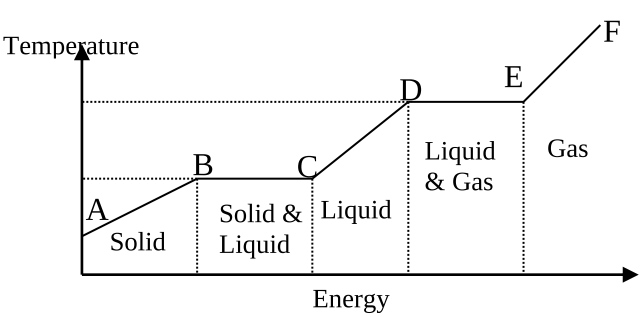
Chapter 9 – Energy & Entropy

- **$\Delta H = \Delta T$ x mass x Cp**; Cp of water is 4.184 J/g °C or 1 cal/ g °C
- The specific heat capacity of a substance is the quantity of heat required to change the temperature of 1 gram of a substance by 1°C.
- Substances with low specific heat capacities are good conductors.
- Substances with high specific heat capacity are good insulators.
- In order for a substance to melt or boil, you must break bonds between molecules.
- Solids are the lowest energy state. Solid substances absorb energy(endothermic) to become liquids and gases. Gases give off energy(exothermic) to become liquids and liquids give off energy to become solids.
- **$\Delta H = H_{fus}$ x mass** (used to calculate the heat needed to melt a substance – watch the units!)
- If heat is a reactant, the reaction is endothermic (ex. . 180 kJ + N_2 + O_2 → 2NO).
- If heat is a product, the reaction is exothermic. (ex. 2NO + O_2 → $2NO_2$ + 112 kJ)





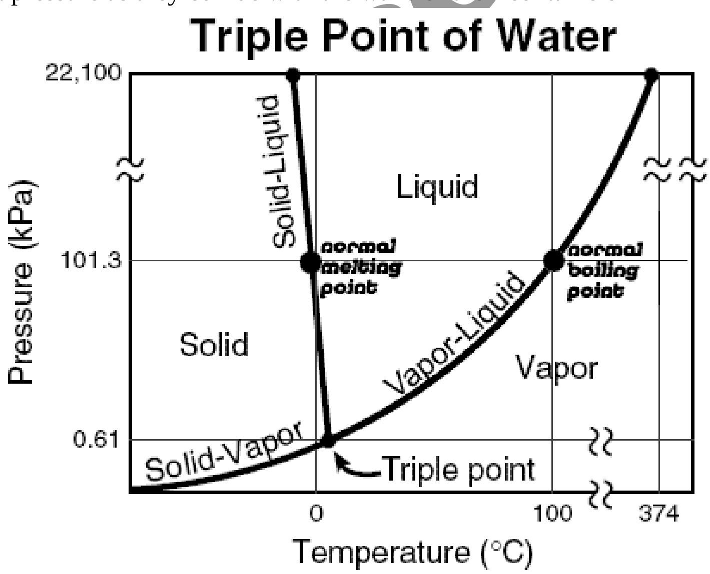
- Catalysts speed up chemical reactions by reducing the amount of activation energy needed to start the reaction. Catalysts are not used up in a reaction.
- Entropy is disorder. Solids have lower entropy than liquids and gases.

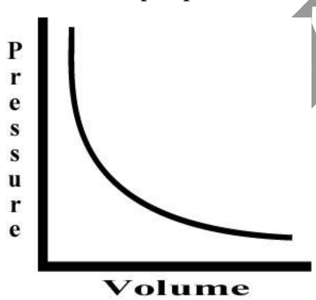


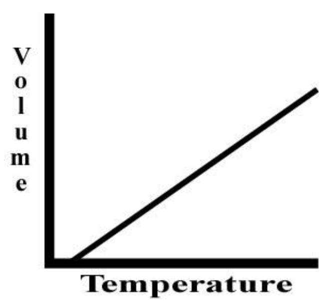
Be familiar with a phase change diagram. During a phase change the heat being added is used to break intermolecular forces.

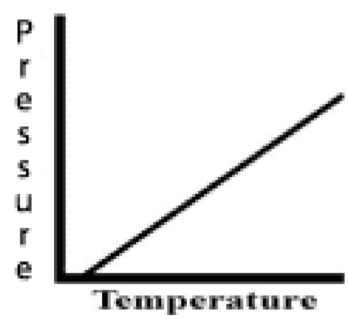
Chapter 10 – Gas Laws

- STP: standard temperature (273 K), and pressure (1.0 atm or 101.3 kPa)
- At STP one mole of a gas has a volume of 22.4 L.
- $P_{total} = P_1 + P_2 + P_3 + \ldots$ Dalton's Law of Partial Pressure (Often used in "collected over water" problems.)
- $P_1V_1=P_2V_2$ Boyle's Law
- $\frac{V1}{T1} = \frac{V2}{T2}$ Charles' Law
- $\frac{P1}{T1} = \frac{P2}{T2}$ Gay-Lussac's Law
- $\frac{P1V1}{T1} = \frac{P2V2}{T2}$ Combined Gas Law
- $PV = nRT$ Ideal Gas Law
- Gases have mass and occupy space.
- Gas particles are in constant motion and exert pressure as they collide with the walls of their containers
- An Ideal Gas does not exist; Real gases have intermolecular forces, particle volume and can change states
- Temperature is a measure of the average kinetic energy of the atoms or molecules. As you increase the temperature, you increase the kinetic energy of the atoms or molecules.
- At the triple point, solid liquid and gas states all exist.
- If you increase the pressure on a liquid you will raise its boiling point. If you reduce the pressure on a liquid, you will lower its boiling point.
- Vapor pressure increases as a substance nears its boiling point.
- At a substance's boiling point vapor pressure is equal to the atmospheric pressure.
- Atmospheric pressure is measured using a barometer.
- Vapor pressure is measured using a manometer.









BOYLE'S LAW
Inverse relationship

CHARLES' LAW
Direct relationship

GAY-LUSSAC'S LAW
Direct Relationship