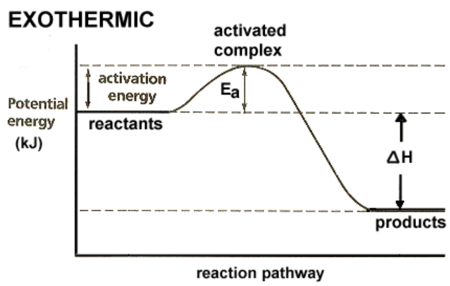
Chemistry Unit 6

Primary reference: *CHEMISTRY*, Addison-Wesley

Topic	Essential Knowledge	Study Support
Atomic Structure 2.6 SOL 2f	**Electronegativity** is the measure of an atom's attraction for electrons in a bond. Electronegativity increases across a period toward the halogens and decreases down a group. The most electronegative atom is fluorine. The least electronegative element (excluding noble gases) is Francium, Fr.	**Ch 14:** Read p. 405
Nomenclature, Formulas, and Reactions 3.6 SOL 3d, 3e, 3f	**Exothermic reactions** release heat whereas **endothermic reactions** absorb heat. **Heat of reaction** is the amount of energy absorbed or released during a chemical change. Exothermic reactions have a negative ΔH_{rxn}, whereas endothermic reactions have a positive ΔH_{rxn}. Examples of writing an exothermic reaction equation are: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_{20} + 890\text{ kJ}$ or $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O,\ \Delta H_{rxn} = -890\text{ kJ/mol}$ **Polar covalent bonds** form between elements with very different electronegativities. The more electronegative atom will attract the electrons more strongly and this will result in it having a slight negative charge. The less electronegative atom then takes on a slight positive charge. A **non-polar covalent** bonds form between atoms of similar electronegativities. A **polar molecule** has unequally distributed electrons around the central atom. This is caused by unsymmetrical polar bonds or a lone pair on the central atom. The positive end of the molecule has a positive dipole and the negative end has a negative dipole. Polar molecules have **dipole-dipole intermolecular attractions** as well as London dispersion intermolecular attractions. Non-polar molecules only have **London dispersion intermolecular attractions**. with O-H, N-H or F-H bonds have **hydrogen bonding** attractions. Kinetics is the study of reaction rates. increase with increased temperature, reactant concentration, increased and the use of a catalyst. **Activation** minimum energy needed to initiate a High activation energies correspond to rates. **Catalysts** speed up reactions  Molecules intermolecular Reaction rates increased surface area **energy** is the reaction. slow reaction rates by decreasing the activation energy. Potential energy diagrams are used to analyze reaction energy changes.	**Ch 11:** Read pp. 303-304 **Ch 16:** Read pp. 460-466 **Ch 19:** Read pp. 533-538.
Molar Relationships 4.6	Stoichiometry can be combined with heat of reaction, ΔH_{rxn}, to calculate the amount of heat produced from a known amount of reactant.	
Phases of Matter and Kinetic Molecular Theory 5.6 SOL 5b, 5c, 5d	Forces of attraction (intermolecular forces) between molecules determine their state of matter at a given temperature. Forces of attraction include hydrogen bonding, dipole-dipole attraction, and London dispersion (van der Waals) forces. **Vapor pressure** is the pressure of the vapor found directly above a liquid in a closed container. When the vapor pressure equals the atmospheric pressure, a liquid boils. **Volatile liquids** have high vapor pressures, weak intermolecular forces, and low boiling points. **Nonvolatile liquids** have low vapor pressures, strong intermolecular forces and high boiling points. **Sublimation** is the phase change from solid to gas without passing through the liquid phase. A substance's **triple point**, is the pressure and temperature conditions where all three phases coexist in dynamic equilibrium. The following mathematical relationship between the pressure, volume and temperature of a gas is used to describe the behavior of gases: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ An Ideal Gas does not exist, but this concept is used to model gas behavior. A Real Gas exists, has intermolecular forces and particle volume, and can change states. The Ideal Gas Law states that $\boxed{PV = nRT}$. R is the ideal gas law constant and has two values depending on the pressure units. They are **R = 8.314 L·kPa/mol·K and R = 0.0821 L·atm/mol·K** **Dalton's Law of Partial Pressures** says the sum of the partial pressures of all the components in a gas mixture equals the total pressure of the gas mixture. $\boxed{P_{total} = P_A + P_B + P_C}$ and $\boxed{n_A/n_{total} = P_A/P_{total} = V_A/V_{total}}$ **Graham's Law** says gas molecules with the lightest mass travel fastest. $\boxed{K.E = 0.5\ mv^2}$	**Ch 10:** Read pp. 269-280 and pp. 284-286. **Ch 12:** Read pp. 350-353.

Unit 6 Objectives
Chemistry, Addison-Wesley, 2002

I) Endothermic and Exothermic Reactions
  A) Classifying Reactions
  B) Stoichiometry and Calculating Heats of Reaction
II) Intermolecular Forces (IMFs)
  A) Polar bonds
  B) Polar molecules
  C) Intermolecular Attractions and Physical Properties
    1) Intermolecular forces
      (a) London Dispersion forces
      (b) Dipole-Dipole attractions
      (c) Intermolecular Hydrogen bonding
    2) Effect of Intermolecular Forces on Physical properties
    3) Comparing molecular and ionic compounds
III) Phase Changes and Intermolecular Forces (IMFs)
  A) Kinetic Energy, Particle Velocity, and Kelvins
  B) Kinetic Energy and Liquids
    1) Vapor pressure
    2) Boiling points and atmospheric pressure
  C) Kinetic Energy and Solids
  D) Phase Changes and Phase Diagrams
IV) Gas Laws: Combined, Ideal, Dalton's Law and Graham's Law
V) Reaction Rates
  A) Collision Theory
  B) Potential Energy Diagrams
    1) Activation Energy
    2) Catalysts

(SOL) Learning Objective

1. (3e) Identify a reaction as endothermic or exothermic based on its thermochemical equation and/or sign of ΔH.
2. (4b) Calculate the heat change of a reaction using stoichiometry and heats of reaction.
3. (2f) Compare the electronegativity of two elements based on their position on the periodic table.
4. (3d)Compare relative bond polarity based on the two elements position on the periodic table.
5. (3d) Use VSEPR theory and electronegativity to identify polar and non-polar molecules.
6. (5d) Identify and compare the three types of intermolecular forces (dipole interaction, hydrogen bonding, London dispersion (van der Waals) forces)
7. (5d) Predict the relative melting and boiling points of molecular and ionic substances based on intermolecular forces.
8. (5d) Explain the relationship between kinetic energy and temperature
9. (5d) Interpret a graph of percent molecules vs kinetic energy
10. (5b) Explain why real gases condense whereas ideal gases do not condense using IMFs and kinetic energy.
11. (5b) interpret vapor pressure graphs.
12. (5d) Explain what happens as a solid melts using IMFs and kinetic energy.
13. (5d) Explain the relationship between a substance's vapor pressure and boiling point and the strength of the substance's IMFs.
14. (5d)Interpret the effect of temperature and pressure on states of matter using a phase diagram.
15. (5d)Identify the triple point on a phase diagram and identify which states of matter exist at the triple point.
16. (5d) Indentify phase changes on a phase diagram of water including fusion, solidification, vaporization, condensation and sublimation.
17. (5b) Solve gas law problems using the Combined Gas Law and the Ideal Gas Law.
18. (5b) Explain the difference between a real gas and an ideal gas.
19. (5b) Predict when a gas will behave most ideally.
20. (5b)Use Dalton's Law to calculate partial pressures
21. (5b) Use Graham's Law to compare rates of effusion and diffusion of two gases
22. (3f) Draw a reaction's potential energy diagram with axes labeled, and ΔH, activation energy, product energy, reactant energy, transition state, and catalyst shift clearly identified for exothermic and endothermic reactions.
23. (3f) Explain how a catalyst increases reaction rate.
24. (3f)Identify and explain the effect the following factors have on the rate of a chemical reaction: (catalyst, temperature, concentration, and reactant particle size).

Chapter 11 Part 2: Endothermic and Exothermic Reactions

Classifying Reactions as Endothermic or Exothermic

Exothermic Reactions ____

Endothermic Reactions____

Heat of Reaction, $\Delta H_{reaction}$ is the heat absorbed or released by a reaction.

+ ΔH_{rxn} _________________________ - ΔH_{rxn}_______________________

Thermochemical Equations ____

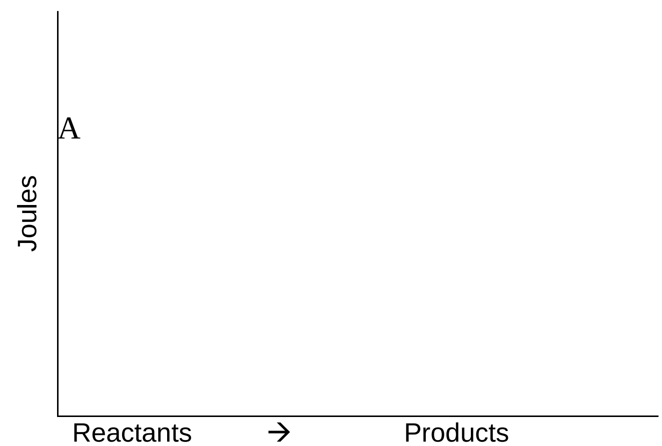
____

There are two general ways they're written. Both are acceptable:

Heat as a Product (exo) or Reactant (endo):
Heat shown as a change in Enthalpy (ΔH) -ΔH means heat is lost (exothermic) +ΔH means heat was absorbed (endothermic)

Example one: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890.4$ kJ so ΔH=______________

Potential Energy Diagram



The system ________________________________ heat so ΔH is negative/positive.

The reaction is endothermic/exothermic?

Law of Conservation of Energy: ____

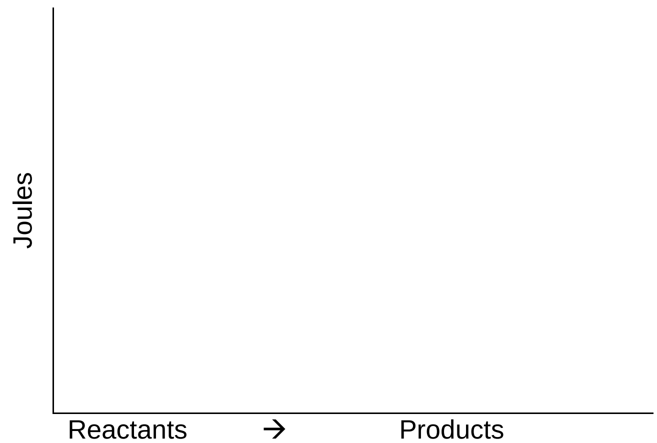
____

Energy = _______________+___________+________________________

Gasoline = ______________________+_____________________+ ____________________

Example Two:

$2NaHCO_3(s) + 129\ kJ \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$ so ΔH=____________



The system ________________________________ heat so ΔH_{rxn} is negative/positive.

C. Heat of Reaction Calculations (Thermal Stoichiometry)

Example 1:

How many kilojoules of energy are produced by burning 821 grams of methane with excess oxygen?(Ans = 45600 J)

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890.4\ kJ$$

How many liters of oxygen would be consumed at STP to produce 122 kJ of heat in the below reaction?(Ans = 6.14 L)

$$890.4\ kJ + CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Solving Heat of Reaction Problems

Treat heat (in J or kJ) the same as any reactant or product in a chemical equation

1. How much heat is produced by the reaction of 25.7 g of CaO in the equation below? (Ans = 29.9 kJ)

$$CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (s) + 65.2\ kJ$$

2. How many grams of $NaHCO_3$ are needed to react completely when 980 kJ of heat are used in the equation below? (Ans = 1300 g)

$$2\ NaHCO_3 + 129\ kJ \rightarrow Na_2CO_3 + H_2O + CO_2$$

3. Using the same equation, how many kJ of heat must be used to produce 55.7 liters of CO_2 at STP?(Ans = 321 kJ)

$$2\ NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2 \qquad \Delta H_{rxn} = +129\ kJ/mol$$

Comparing Endothermic and Exothermic Reactions

Exothermic Reactions	**Endothermic Reactions**
The reaction ____________________ heat Example of exothermic reaction	The reaction ____________________ heat Example of endothermic reaction
Potential Energy vs. Reaction Path for Exothermic Reactions	Potential Energy vs. Reaction Path for Endothermic Reactions
An exothermic reaction has a net __________________ ΔH	An endothermic reaction has a net _____________________ΔH
In exothermic reactions, the product's energy is ______________ than the reactant's energy	In endothermic reactions, the product's energy is _______________ than the reactant's energy

Word bank: positive, releases, negative, absorbs, lower, higher

Chapter 16: Polar Bonds and Polar Molecules and Intermolecular Forces

Electronegativity:____

____

A. Non-Polar and Polar Covalent Bonds

1) non-polar covalent bonds: atoms share bonding electrons equally.

   Example: Cl_2

2) polar covalent bonds (polar bonds): bonding electrons shared unequally.

   Example: HCl

3) Electron sharing based on electronegativity differences.
   a) more electronegative atom attracts the electrons more closely and acquires a slight negative charge.
   b) less electronegative atom then acquires a slight positive charge.
   c) unequal sharing creates “polarized” bonds with opposite charges.
   d) Two ways to show polarity in structural formulas.

   lower case greek deltas: ____

   slashed arrows : ____

The type of bond depends on electronegativity differences between the atoms

Electronegativity Difference	Guideline: Type of Bond	Example (electronegativity difference)
0.0 – 0.4	Non-polar Covalent	C-H in CH_4
0.4 – 2.0	Polar Covalent	HF
>2.0	Ionic	NaCl

Selected Electronegativity Values

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8

Is the bond polar, non-polar or ionic?

C=O in CO_2?
Si-H in SiH_4
C-F in CF_4
N-Cl in NCl_3
KCl

Answer this question without looking at the table on the previous page.

Which bond is most polar; C-N or C-F?

B: POLAR MOLECULES

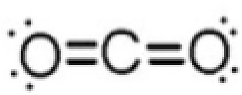
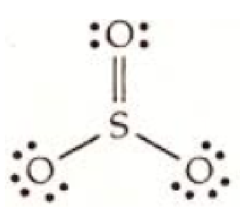
Polar Molecules: One end of the molecule is slightly negative, and one end is slightly positive.

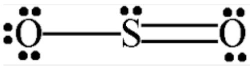
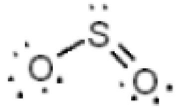
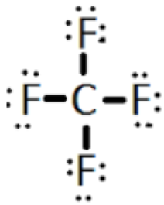
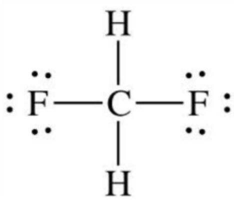
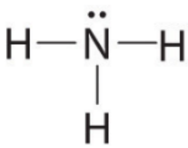
Dipole: a molecule with two poles (one negative, one positive or δ-, δ+)

What makes a molecule polar?	A molecule is polar if the electrons are pulled to one side of the molecule. The molecule is lopsided (assymetrical).

Determining if a molecule is polar.

1. Draw the Lewis structure
2. Determine the molecular geometry
3. Look for lone pairs on central atom (automatically polar)
4. Are there polar bonds?
5. If yes, are the polar bonds unsymmetrical in 3-D around the molecule's center?

Molecule	Lewis Structure and Geometry	Polar or Non Polar
CO	**:C≡O:**	
CO_2		
SO_3		

Molecule	Lewis Structure and Geometry	Polar or Non Polar
SO_2	 	
CF_4		
CH_2F_2		
NH_3		
H_2O		

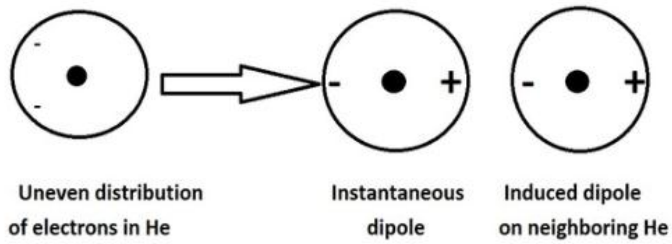
Which molecule is most polar: HCl or HI?

Which molecule is most polar; CO_2 or SO_2?

C: INTERMOLECULAR FORCES

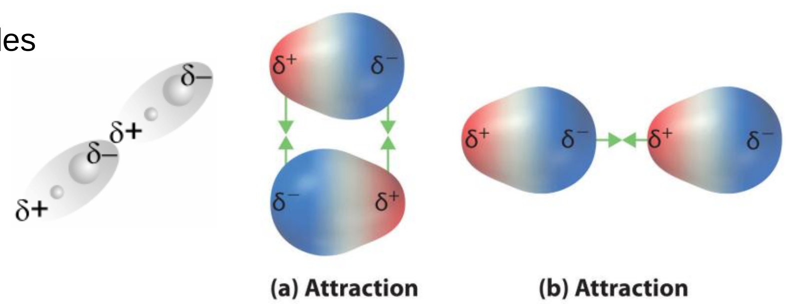
<u>Intermolecular Forces</u> are attractions between molecules due to three forces

1. **<u>London Dispersion Forces</u>** (weakest) **temporary** attractions between molecules due to temporary dipoles caused by shifting electron clouds. Dispersion forces are greater in more massive molecules with larger electron "clouds". **All molecules have LDFs**, and **larger atoms/molecules have stronger LDF because they have more electrons**.

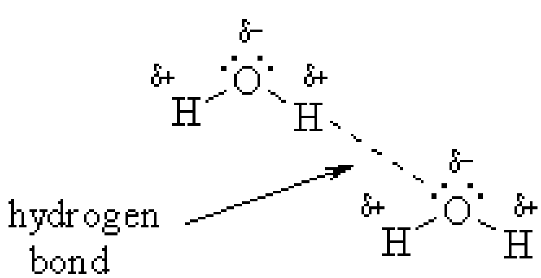


2. **<u>Dipole –dipole attraction</u>**: **polar** molecules are attracted to each other (*like magnets*). The positive dipole of one molecule is attracted to the negative dipole of another. These occur when molecules have uneven distribution of electrons due to electronegative elements (like N, O, Cl, F, Br, etc.).

Example: HCl molecules



3. **<u>Intermolecular Hydrogen bonding</u>**: hydrogen that is *covalently* bonded to a **<u>very</u>** electronegative atom is also weakly bonded to the unshared pair of another electronegative atom.
<u>H</u>-Bonding only occurs when an molecule has H-____, H-____, or H-____ bonds.
Why *these* three atoms specifically?



4. Intermolecular Forces and Molecular Physical Properties

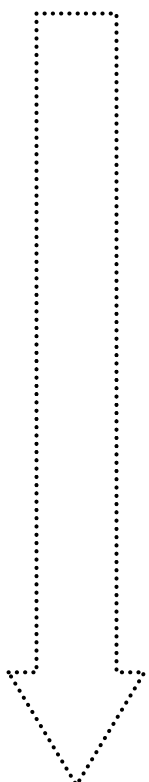
As intermolecular forces (IMF) increases (meaning gets stronger), the melting and boiling points increase because more kinetic energy is needed to overcome the IMFs between molecules.

An Analysis of the Halogens and their Physical States:

The Halogens are **Di**atomic:
(Lewis diagram→)

Are pure diatomic halogen elements polar? Explain & draw dipole vectors & partial charges.

What is the only IMF that all non-polar things can do?



Halogen (X)	Molar Mass (X_2)	State at Room Temp.

Comparing Ionic and Covalent Compounds

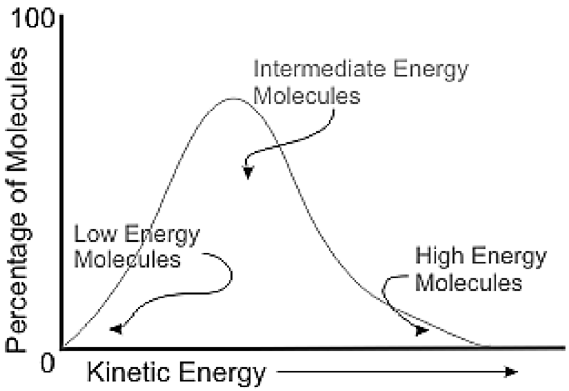
Characteristic	Ionic Compound	Molecular Compound
bond formation		
Types of elements in compound		
Physical state at 25˚C		
Melting point		
Electrical Conductivity in aqueous solution		

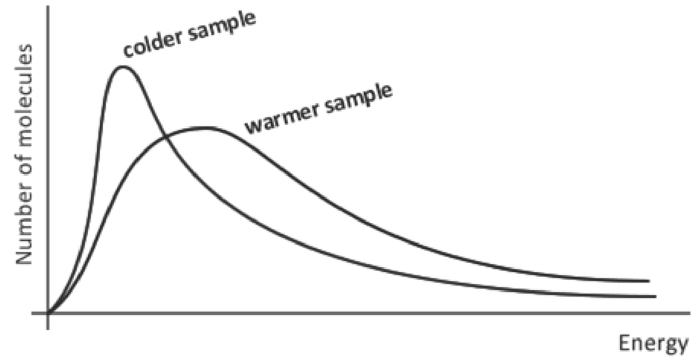
Chapter 10 Kinetic Theory, IMFs, and Phase Changes

Kinetic Molecular Theory: The tiny particles in all forms of matter are in constant ______________.

I Kinetic Energy and Kelvin temperature scale

A) Temperature measures average kinetic energy





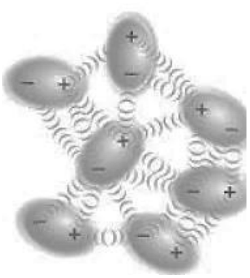
B) Gas particle's kinetic energy increases as ____

C) Kelvin Temperature scale is ____

____________K = __________°C

II.Kinetic Energy and Liquids

Intermolecular forces (between molecules) hold particles together in solid or liquid phases.



Kinetic energy keeps the molecules moving but not with enough energy to overcome the IMFs.

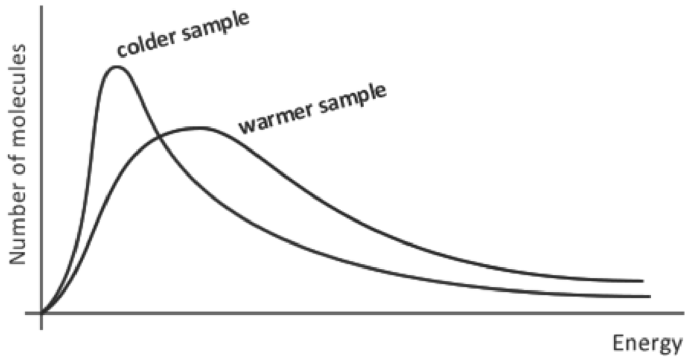
Evaporation, Vapor Pressure and Temperature

Evaporation:____

____

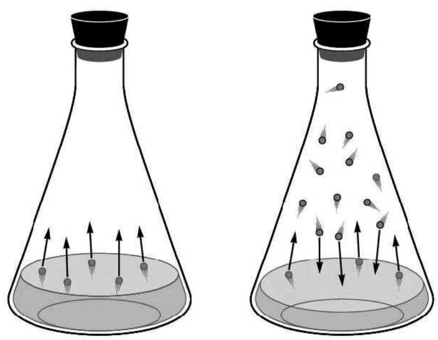
____

- Particles with enough kinetic energy to overcome intermolecular forces escape into gas phase
- Evaporation rate increases as temperature increases



Evaporation in a closed container produces <u>vapor pressure.</u>

Increasing temperature increases vapor pressure over a liquid until a <u>dynamic equilibrium</u> is reached.



Volatile liquids____

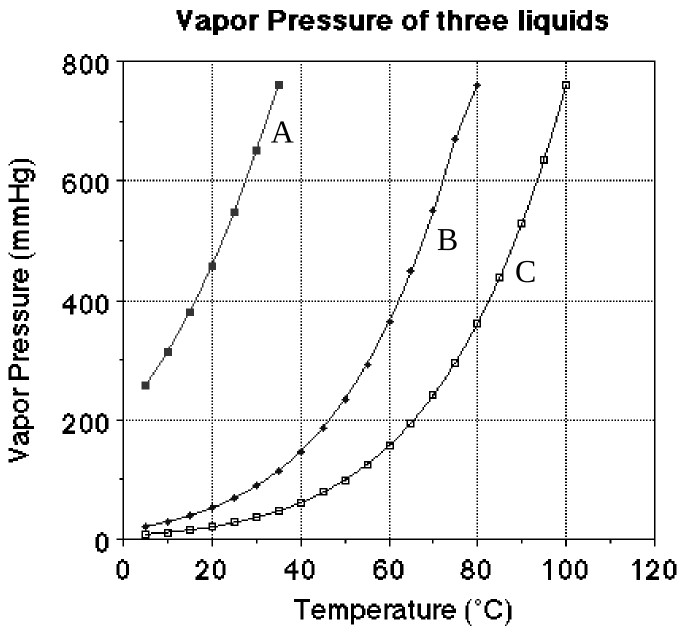
Nonvolatile liquids ____

Boiling Point of a Liquid (open container)

Boiling Point:____

____

A) Boiling point changes as external pressure changes/



What is the vapor pressure of liquid A at 20°C?_

Which liquid represents water?

What is the boiling point of B when the external pressure is 400 mmHg?

Which liquid is most volatile?

Which liquid has the strongest intermolecular forces?

How hot does water need to be to boil at 100 mmHg?

What will be the boiling point of water on Pike's Peak (elevation = 14,000 ft, atmospheric pressure = 640 mmHg)?

III.Kinetic Energy and Solids

melting point:____

____

____

sublimation:____

Examples:

$CO_2(s) \rightarrow CO_2(g)$

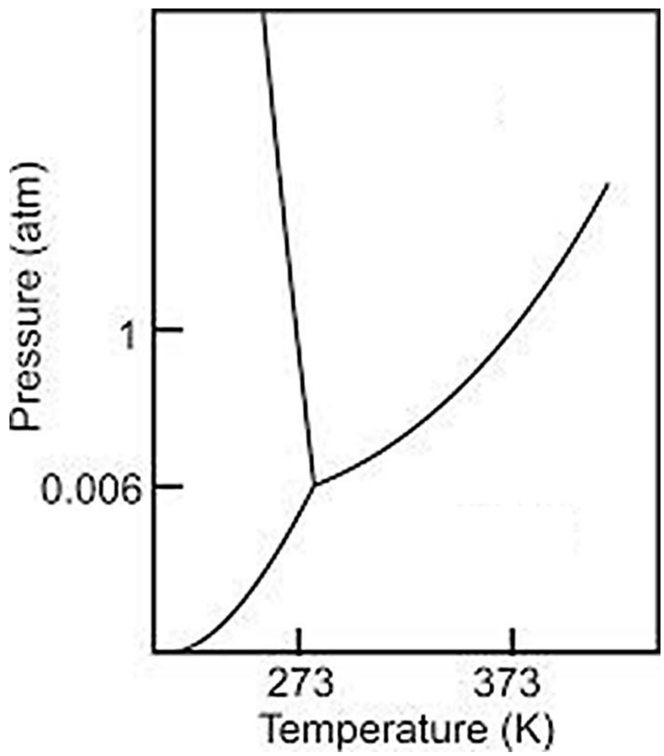
$I_2(s) \rightarrow I_2(g)$

IV. Phase Changes and Phase Diagrams

Phase Changes and terms

SOLID **LIQUID** **GAS**

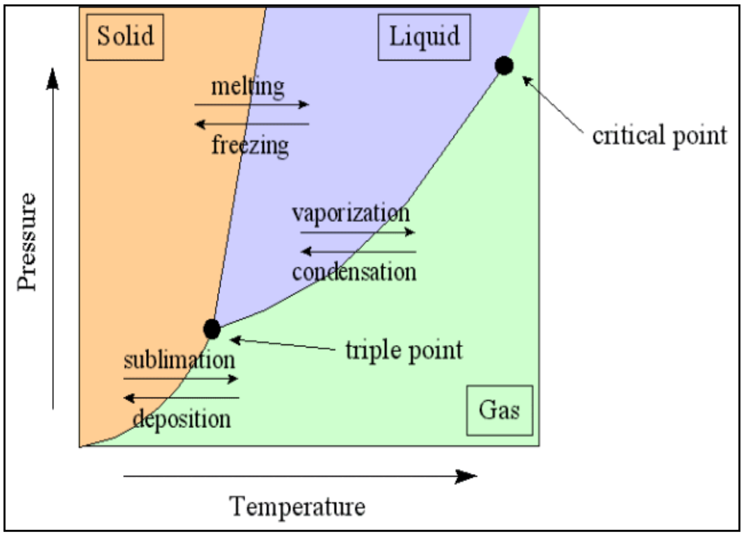
Phase Diagram and Triple Point



Phase Diagrams

A **phase diagram** is a graphical way to depict the effects of **pressure** and **temperature** on the phase of a substance:

The CURVES indicate the conditions of temperature and pressure under which "equilibrium" between different phases of a substance can exist. BOTH phases exist on these lines:



Melting/Freezing: Any point on this line (pressure & temperature) the substance is both **solid** and **liquid**

Sublimation/Deposition: Any point on this line (pressure & temperature) the substance is both **solid** and **gas**

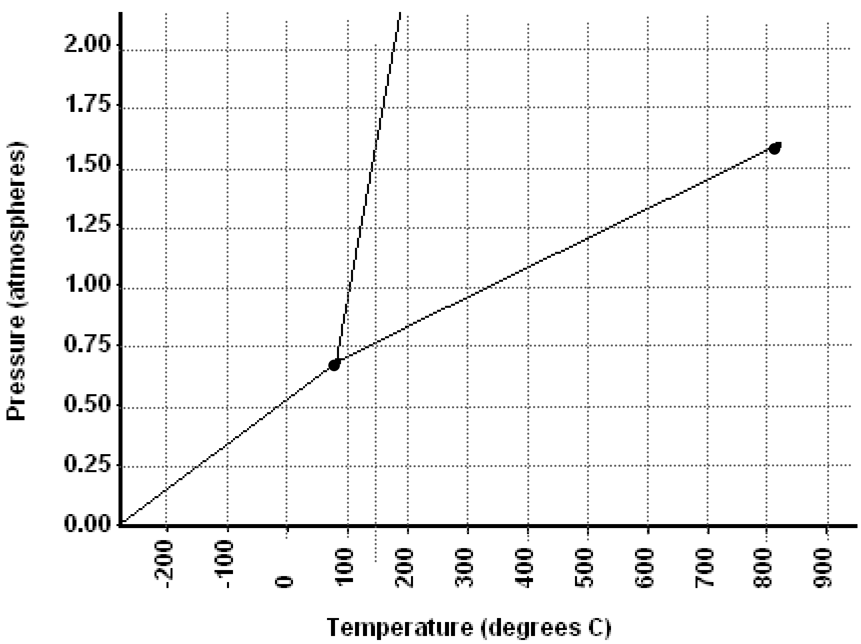
Vaporization/Condensation: Any point on this line (pressure & temperature) the substance is both **liquid** and **gas**

NOTE: the vapor pressure curve ends at the **critical point,** the temperature above which the gas cannot be liquefied no matter how much pressure is applied (the kinetic energy simply is too great for attractive forces to overcome). Any substance beyond this critical point is called a **supercritical fluid** – *indistinguishable* between gas or liquid (*neither one*)

The **TRIPLE POINT** is the condition of temperature and pressure where ALL THREE phases exist in equilibrium (solid, liquid, gas)

Remember that pressure can be expressed in many units where: **1 atm = 101.3 kpa = 760 mmHg = 760 torr = 14.7 psi**

Refer to the phase diagram below when answering the questions.
NOTE: "*Normal*" refers to STP – Standard Temperature and Pressure.



1) What are the values for temperature and pressure at STP? T= ________, P= ________

2) What is the *normal* freezing point of this substance? ________

3) What is the *normal* boiling point of this substance? ________

4) What is the *normal* melting point of this substance? ________

5) What is the phase (s, l, g) of a substance at **2.0 atm** and 100 °C? _________

6) What is the phase (s, l, g) of a substance at **0.75 atm** and 100 °C? _________

7) What is the phase (s, l, g) of a substance at **0.5 atm** and 100 °C? _________

8) What is the phase (s, l, g) of a substance at 1.5 atm and **50 °C**? _________

9) What is the phase (s, l, g) of a substance at 1.5 atm and **200 °C**? _________

10) What is the phase (s, l, g) of a substance at 1.5 atm and **800 °C**? _________

11) What is the condition of the **triple point** of this substance? T= ________, P= _______

12) If a quantity of this substance was at an initial pressure of 1.25 atm and a temperature of **300^0 C** was lowered to a pressure of 0.25 atm, what phase transition(s) would occur? ____________

13) If a quantity of this substance was at an initial pressure of 1.25 atm and a temperature of **0^0 C** was lowered to a pressure of 0.25 atm, what phase transition(s) would occur? ____________

14) If a quantity of this substance was at an initial pressure of **1.0 atm** and a temperature of 200^0 C was lowered to a temperature of -**200^0 C**, what phase transition(s) would occur? ____________

15) If a quantity of this substance was at an initial pressure of **0.5 atm** and a temperature of 200^0 C was lowered to a temperature of -**200^0 C**, what phase transition(s) would occur? ____________

16) If this substance was at a pressure of 2.0 atm, at what temperature would it **melt**? ____________

17) If this substance was at a pressure of 2.0 atm, at what temperature would it **boil**? ___________

18) If this substance was at a pressure of 0.75 atm, at what temperature would it **melt**? ___________

19) If this substance was at a pressure of 0.75 atm, at what temperature would it **boil**? ___________

20) At what temperature do the gas and liquid phases become indistinguishable from each other? ______

21) At what pressure would it be possible to find this substance in the gas, liquid, **and** solid phase? _____

22) If I had a quantity of this substance at a pressure of 1.00 atm and a temperature of -100^0 C, what phase change(s) would occur if I **increased the temperature** to 600^0 C? At what temperature(s) would they occur? (***NOTE***: *multiple answers needed for this question*)

22) If I had a quantity of this substance at a pressure of 2.00 atm and a temperature of -150^0 C, what phase change(s) would occur if I **decreased the pressure** to 0.25 atm? At what pressure(s) would they occur? (***NOTE***: *multiple answers needed for this question*)

Chapter 12 Dalton's Law and Graham's Law

Mixtures of Gases: Dalton's Law of Partial Pressure

Partial Pressure: ____

____

Verbally: At constant pressure and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of the component gases.

Math Equation:

Example 1: What is the partial pressure of oxygen in air at STP (101.3 kPa) if P_{N2} = 79.1 kPa, P_{CO2} = 0.040 kPa and P_{others} = 0.94 kPa?

Example 2: A sample of oxygen gas is collected over water at 20.°C. The vapor pressure of water at 20.°C is 15 mm Hg. If the total pressure is 420 mm Hg, what is the partial pressure of the oxygen?

Variants of Dalton's Law

$$\frac{mole_A}{mole_{Total}} = \frac{P_A}{P_{Total}} = \frac{\%V_A}{100\ \%V_{Total}}$$

Example 1: A tank contains 6.0 moles of a mixture of hydrogen, helium, and nitrogen at 102 kPa. If there are 2.0 moles of hydrogen in the tank, what is the partial pressure of hydrogen?

Example 2: A gas cylinder contains 8.0 moles of argon, 2.0 moles of nitrogen, and 2.0 moles of oxygen at 600. mmHg. What is the partial pressure of nitrogen in the cylinder?(Ans = 1.0×10^2 mmHg)

Example 3: A mixture of gases with a pressure of 950 mm Hg contains 20% hydrogen and 80% neon by volume. What is the partial pressure of neon gas in the mixture?(Ans = 760 mmHg)

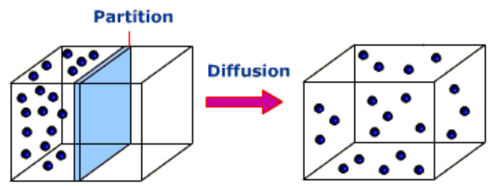
Example 4: In a mixture of oxygen and nitrogen gas, 70.0 percent of the total gas pressure is exerted by the nitrogen. If the total pressure is 150 kPa, what pressure does oxygen exert?(Ans = 45 kPa)

Graham's Law of Effusion

Diffusion ____

____

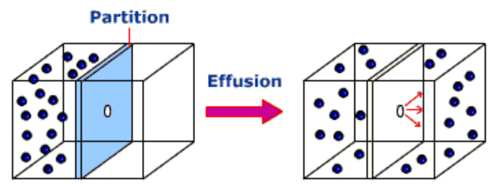
____



Effusion____

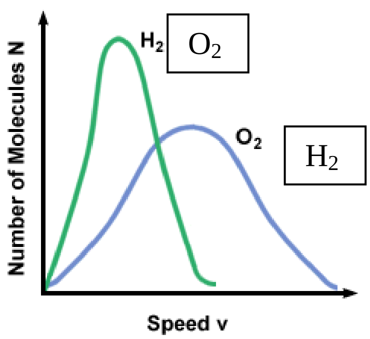
____

____



Graham's Law Verbally: Gas molecules with the lightest mass travel fastest.

Graham's Law Equation:



Which gas will escape slowest from a tiny hole in a balloon; He, C_3H_8, or Xe?

Partial Pressures Practice (Lee) NDB→

1) **Argon, oxygen, and nitrogen are mixed together and pressurized in a tank. P_{O2} is 155 kPa, P_{N2} is 415 kPa, and P_{Ar} is 285 kPa.**
    a. What is the total pressure of the gas mixture?

    b. What is the % O2?

    c. What is the % N2?

2) **The total pressure in a tank is 1200.0 mmHg. Krypton has a partial pressure (P_{Kr}) of 680.0 mmHg, and methane has a partial pressure that is half that of krypton. The third gas in the container is chlorine.**
    a. Calculate the partial pressures of each gas with *units*:

       P_{CH4} ____________________

       P_{Kr} _____________________

       P_{Cl} _____________________

    b. Determine the % volume of each gas:

       % Cl ____________________

       % Kr ____________________

       % CH_4 ____________________

3) **A highly pressurized (46.2 atm) mixture of gases contains a total of 330. moles of gases. The technologist who created the mixture added fluorine, chlorine, helium, and hydrogen. Twice as many moles of helium were present than fluorine. Three times as many moles of chlorine were added than fluorine, and the half as many moles of hydrogen were added as fluorine.**
    a. Set up an algebraic equation using variables that expresses the situation above.

    b. Determine the number of moles of each gas.

    c. Determine the partial pressures of each gas. Use proper notation. (Example: To express the partial pressure of X, write P_X)

4) A certain planet was discovered whose atmospheric composition was 70% CO_2, 20% O_2, and 7% H_2, and 3% He. The atmospheric air pressure on the planet was determined to be twice that of normal atmospheric pressure on Earth. Calculate the partial pressure of each gas *<u>in kPa</u>*.

5) 15.0 g of nitrogen gas, and 15.0 grams of chlorine gas were added to a container that exists at STP.

a. What is the number of moles of nitrogen in the container? ____________________

b. What is the number of moles of chlorine in the container? ____________________

c. How many total moles of gas are in the container? ___________________

d. What is the P_{N2} in kPa? ___________________

e. What is the P_{Cl2} in kPa? ___________________

Graham's Law of Effusion (Applied Quantitative Practice)

Graham's Law can be derived from the equation for kinetic energy (K =)

Derivation:

Graham's Law (final equation) starts that lighter gases move more quickly than heavier gases in an *inverse square* proportion. (According to kinetic theory, when the same amount of energy is available to different bodies with different masses, they will move at inverse square velocities relative to each other.) A fundamental assumption when using Graham's Law of Effusion is that the gases have the **<u>same amount of energy</u>**… so they're at the **<u>same temperature</u>** on the Kelvin scale. Their difference in average velocity (aka: rate, speed…) is due to mass difference.
Analogy: A 60-kg girl eats 2 eggs and 3 slices of bacon. A 150 kg sumo wrestler eats the same thing. The girl will run faster because she's smaller, even though they had the same breakfast (energy).

Graham's Law of Effusion

Tips for Use:

Worked Example 1: The average velocity of oxygen (O_2) molecules will be faster than the average velocity of chlorine (Cl_2) molecules because oxygen has a smaller molar mass. What is the relative rate (speed) of oxygen molecules to chlorine molecules if they are at the same temperature? (i.e., how many times faster will molecules of oxygen move?)

Part 1: Determine which gas will effuse or move fastest, and determine how many times faster it moves.

1) H_2 vs. He

2) O_2 vs. Ne

3) CH_4 vs. NCl_3

4) Ammonia vs. Hydrogen Sulfide (H_2S)

5) Xenon vs. Argon

Part 2: Determine which gas effuse or move slowest, and determine how many times slower it moves.

6) Oxygen vs. chlorine

7) Sulfur dioxide vs. methane

8) Laughing gas (dinitrogen monoxide) vs. carbon monoxide

9) Sulfur Hexafluoride vs. carbon tetrafluoride

10) Silane (SiH_4) vs. hydrogen

Part 3 (Advanced): Determining the molar mass of an unknown gas; or identifying the gas by calculating the molar mass from relative rates.

11) A sample of hydrogen gas effuse through a porous container 9 times faster than an unknown gas. Estimate the molar mass of the unknown gas. Would is reasonable to assume this gas is silicon tetrafluoride? Explain and justify your answer using mathematics and complete sentences.

12) At a certain temperature, hydrogen molecules move at an average velocity of 1.84×10^3 m/s. Estimate the molar mass of a gas whose molecules have an average velocity of 311 m/s

13) Nitrogen gas (N_2) effuses at a rate 2.17 faster than an unknown noble gas. Identify the noble gas

14) A sample of $Br_2(g)$ take 10.0 min to effuse from one side of a room that is 86 feet long. How long would it take the same amount of Ar(g) to effuse the same distance?

15) Explain why carbon monoxide and nitrogen effuse at nearly the same rate. Use complete sentences and justify your answer.

Chapter 19 Part 1: Reaction Rates

http://www.wwnorton.com/college/chemistry/gilbert/index/site_map.htm

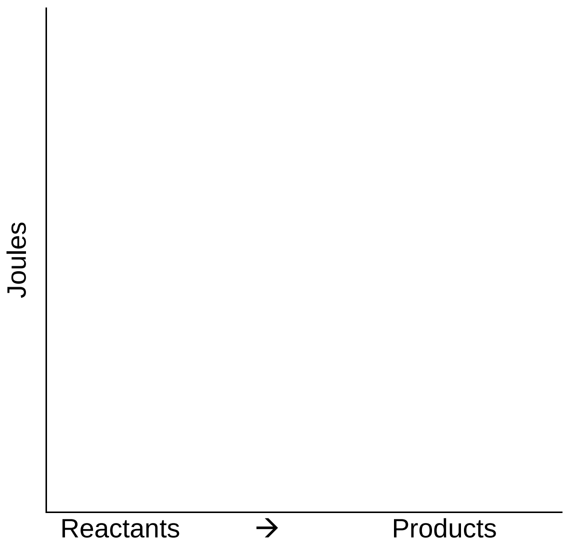
Reaction rates are measured as mol/time units.

A. Collision Theory:____

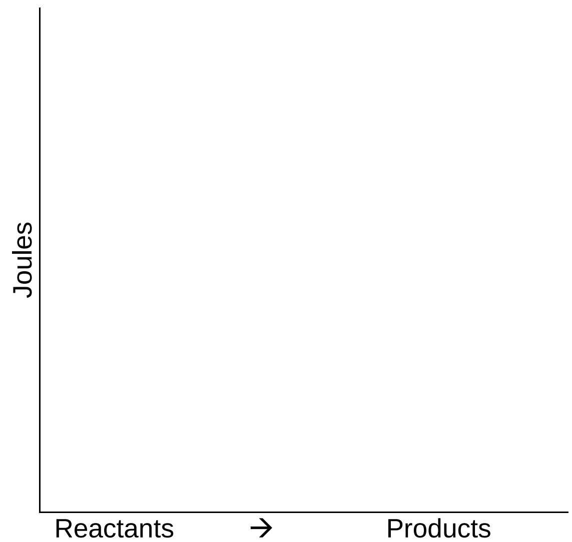
____

Reaction Coordinate Diagram for Exothermic Process (forward direction)

Chapter 11: Thermochemistry



Collision Theory



Activation Energy:____

____

Activation Complex:____

____

B. Factors Influencing Reaction Rate

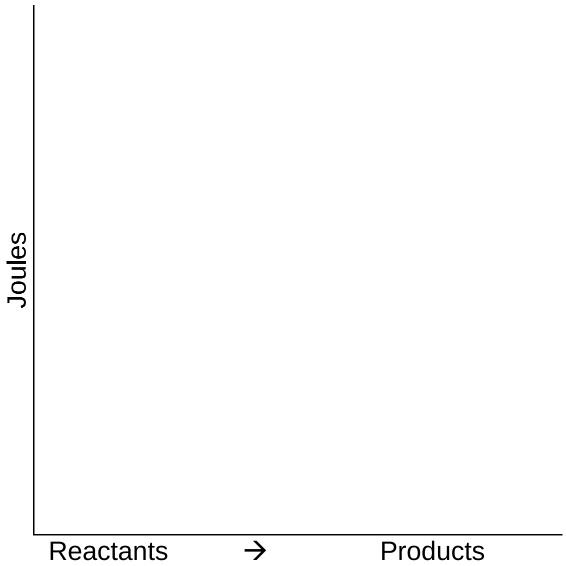
1. Temperature

2. Concentration

3. Particle Size

4. Catalysts

Catalyst Effect on Reaction Path

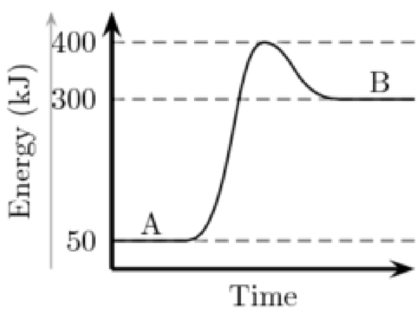


Why do catalysts increase reaction rate?

Examples: Catalytic Convertors

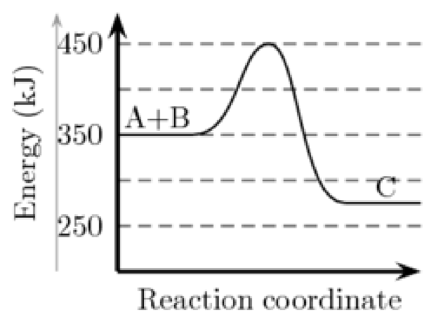
Interpret the following potential energy diagrams

Reaction 1



A → B

Reaction 2



A + B → c

1. Which reaction is endothermic?_____________
2. What is the activation energy of Reaction 1? ________
3. What is the ΔH_{rxn} of reaction 1?___________
4. What is the activation energy of reaction 2?_________
5. What is the ΔH_{rxn} of reaction 2?_____________
6. Sketch the effect of a catalyst on both reactions
7. Does a catalyst effect the ΔH_{rxn}?

Biologically (in terms of **biochemistry**), ________________________ **are catalysts.**

How are they usually used? ____

____

________________________ is the process by which life is ________________________.

(meaning "to build up")

(meaning "to break down")

Anabolic ("constructive") processes are when larger, more complex biomolecules are created from _______________ ones.

Example #1: When your body is building _______________ proteins using _________________ _________________.

Example #2: **PHOTOSYNTHESIS** Plants use ________ & ________ to create _____________________________, which is a larger molecule.

Anabolism = Energy ______________

Catabolic ("destructive") processes are when ____________, ____________ molecules are broken down into smaller ones biomolecules.

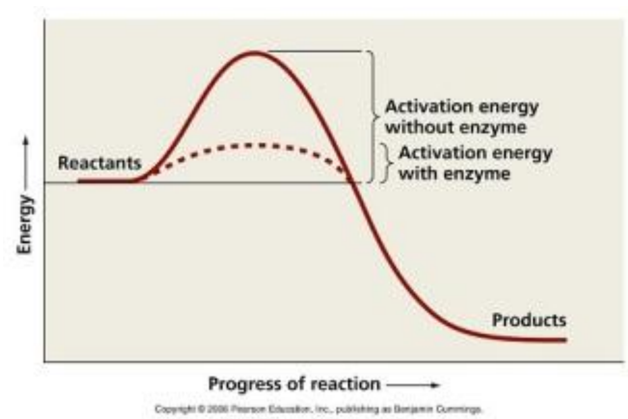
Example #1: **RESPIRATION** & **GLYCOLYSIS**: Your body breaks down ______________________/ ___________ to produce energy, ________ & ________.

Example #2: When you exercise, fats are ____________ ____________.

Catabolism = Energy ____________

Enzymes (which are biological _________________) speed up reactions, but they are NOT ____________________ in the reaction.

They are ________________, which means they're made of amino acids.



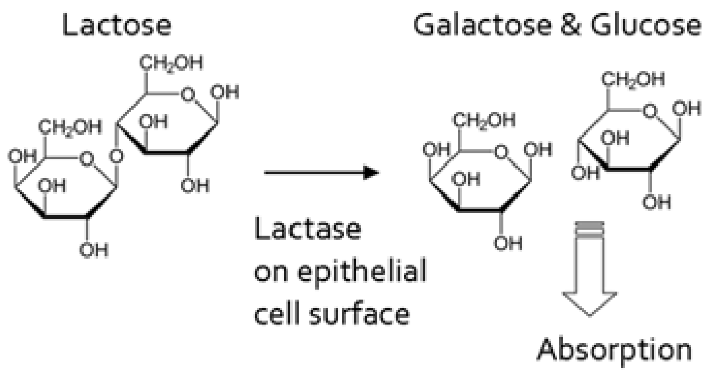
Enzymes usually end in the suffix ____________.

Carbohydrates often end in ____________.

Lactose Intolerance:

Lactose is a ______________________ that is found naturally in _________________ products.

People are who lactose intolerant have a difficult time breaking down the lactose molecule.

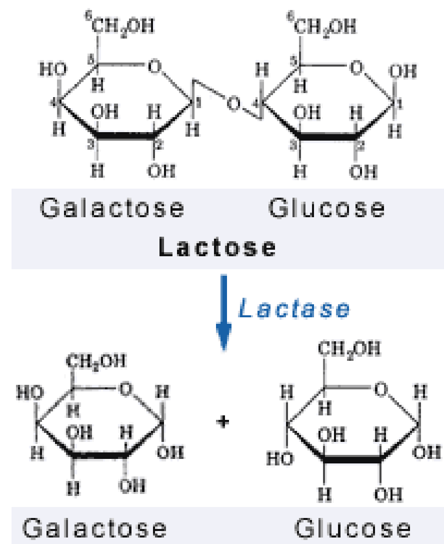


What suffix do carbohydrates often have?

If you see dextrose, maltose, or sucrose on a food label, would you call them fat? Sugar? Or protein?

Lactose is a 2-ring sugar. It must be __________________ by the enzyme ________________.

This will turn it into ______ _____________ ______________, which the body can then use.



People who are *lactose intolerant* don't have enough lact***ase***

enzymes in their ______________ to break down the lact***ose*** sugar.

THINK: What kinds of problems result from consuming lactose with insufficient or non-existent gut lactase?

Other Enzymes - Research their bodily functions on your own:

DNA Helic***ase*** ____

DNA Polymer***ase*** ____

Amyl***ase*** ____

Prote***ase*** ("PRO-tee-ase")____

Lip***ase*** ____