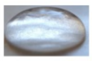
# Solids-Liquids
# Topic#12
# AMSAT Chem 1H

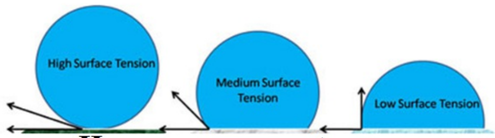
Student Edition

# Solid-Liquids
# Topic#12

## Liquids

1) KE($l$) < KE($g$)
2) Particles are touching and held loosely by intermolecular forces (IMF)
   - LDF/dipole-dipole/hydrogen-bonds)
3) **Intermolecular forces** are weak enough to allow for movement of particles past one another. Fluid and flow.
4) Constant, random motion, slower than gases.
5) More ordered than gases.
6) High density.
7) Incompressible.
8) Diffusion, but slower than gases.



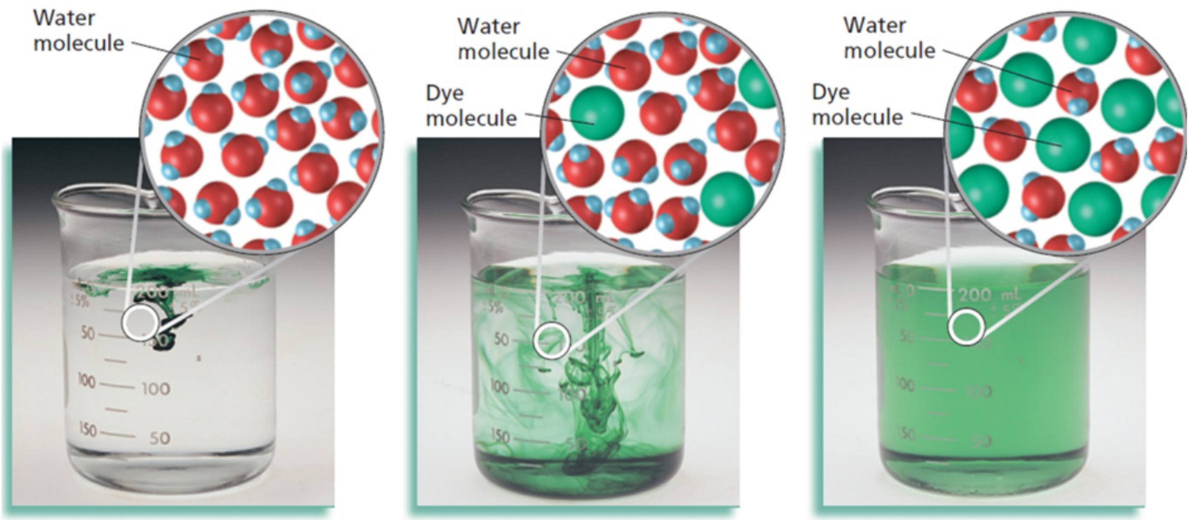



Hg water
ethanol

9) **Surface tension** - force sideways, and below liquid surface between particles at the surface and particles below. Cohesive force "inside liquid."
10) **Capillary action** - adhesion between liquid and solid surface of a tube and cohesion between particles in the liquid.
11) **Evaporation** - liquid particles become gas particles below boiling point and above freezing point.
12) **Boiling point** (*bp*) - point where *VP* above liquid is EQUAL to the atmospheric (ambient) air pressure. Bubbles of vapor form in liquid.
13) **Freezing point** (*fp*) - solidification, removal of energy, as heat, from a liquid.
14) **Viscosity** - the internal friction between parts in a liquid; resistance to flow. A measure of the strength of the intermolecular forces holding a liquid together.

## Solid-Liquids
## Topic#12

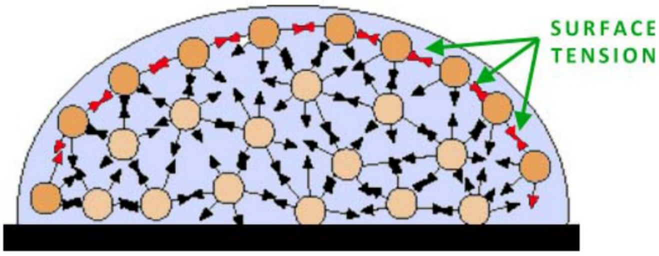
**Liquids**

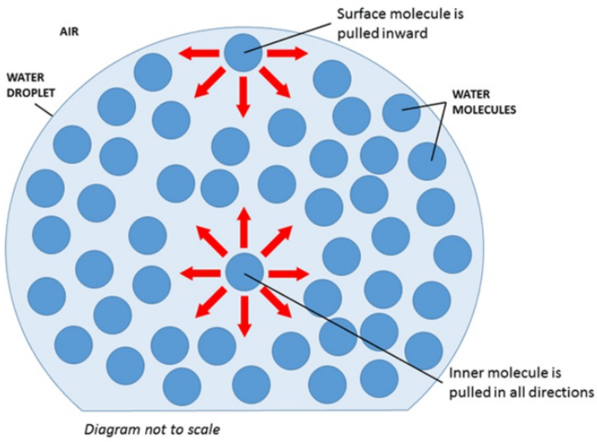


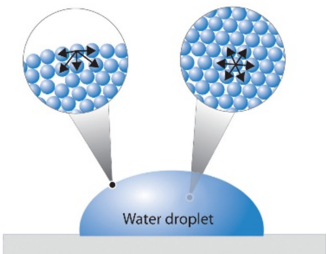

Like gases, the two liquids in this beaker diffuse over time. The green liquid food coloring from the drop will eventually form a uniform solution with the water.

# Solid-Liquids
# Topic#12

## Liquids

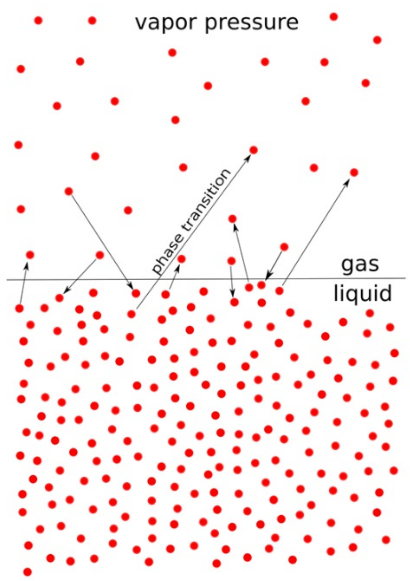



# Solid-Liquids
# Topic#12

## Liquids

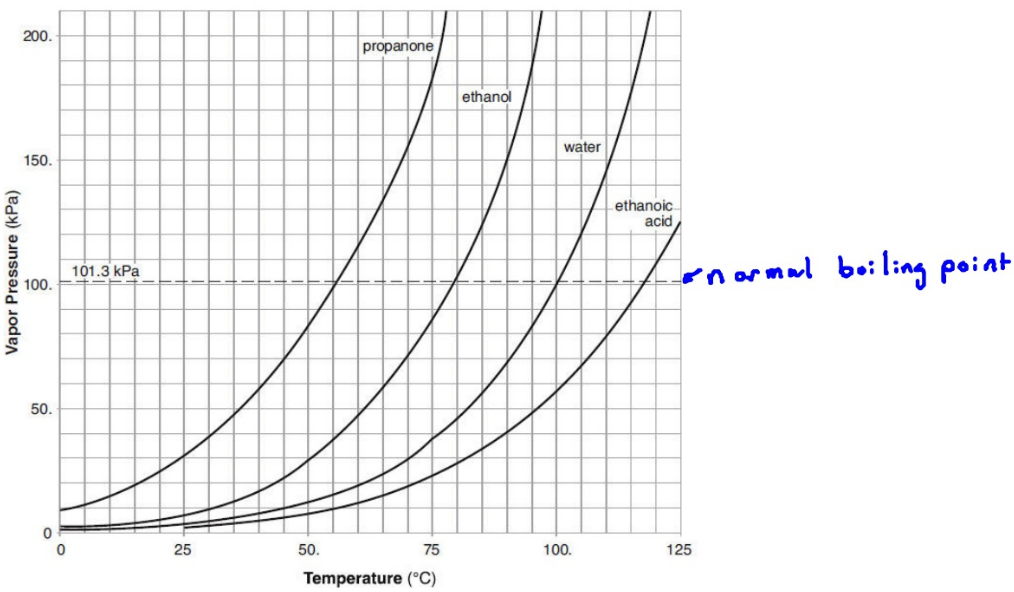



| Vapor Pressure of Water | | | | | |
|---|---|---|---|---|---|
| Temperature (°C) | Pressure (kPa) | Temperature (°C) | Pressure (kPa) | Temperature (°C) | Pressure (kPa) |
| 0 | 0.6 | 21 | 2.5 | 30 | 4.2 |
| 5 | 0.9 | 22 | 2.6 | 35 | 5.6 |
| 8 | 1.1 | 23 | 2.8 | 40 | 7.4 |
| 10 | 1.2 | 24 | 3.0 | 50 | 12.3 |
| 12 | 1.4 | 25 | 3.2 | 60 | 19.9 |
| 14 | 1.6 | 26 | 3.4 | 70 | 31.2 |
| 16 | 1.8 | 27 | 3.6 | 80 | 47.3 |
| 18 | 2.1 | 28 | 3.8 | 90 | 70.1 |
| 20 | 2.3 | 29 | 4.0 | 100 | 101.3 |

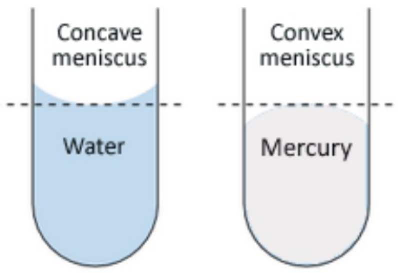
## Solid-Liquids
## Topic#12

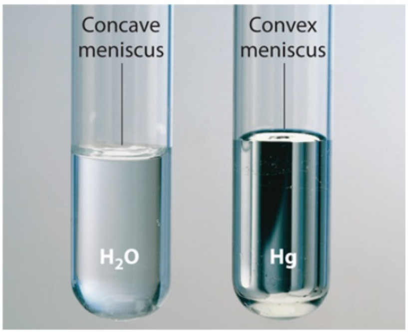
**Liquids**

Table H
Vapor Pressure of Four Liquids



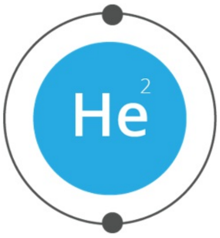
## Liquids


Concave meniscus
Convex meniscus
Water
Mercury

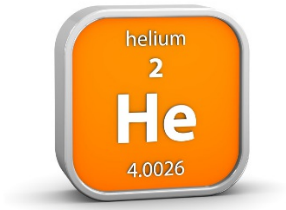

Concave meniscus
Convex meniscus
$H_2O$
Hg

## Solid-Liquids
## Topic#12

**Liquids**




**Super Fluid!**




**Super Fluid!**

**The Magic of Ferrofluid**

## Solid-Liquids
## Topic#12

**Liquids**

Gallium - The Terminator Metal


The Leidenfrost Effect

Non-Newtonian Fluids (1)

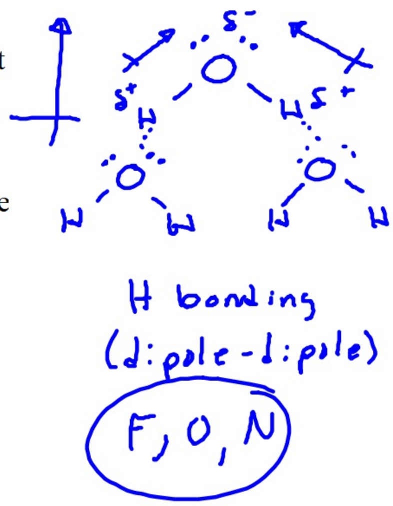
Non-Newtonian Fluids (2)

**Solid-Liquids**
**Topic#12**

## Water

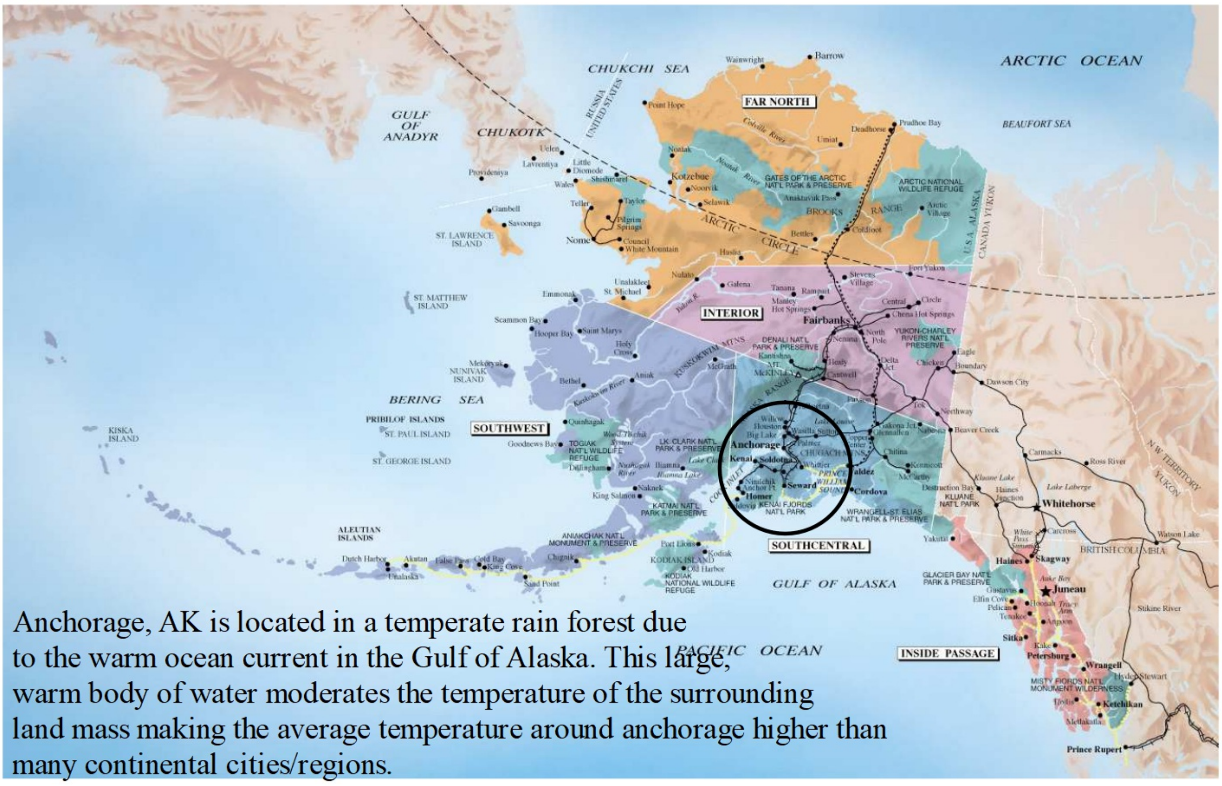
**Water, what is so special about it, anyway?**

- You are about 60% water
- Unique properties stem from the strong intermolecular hydrogen bonds that form between polar water molecules.
- High boiling point for a molecular compound ($NH_3$, $H_2S$, and HF are corrosive gases at room temperature.
- Can absorb/release large quantities of heat. Large bodies of water moderate the surrounding temperatures. Anchorage, Alaska is a moderate rain forest.
- The structure of water expands when turning into a solid. The density of liquid water is $1.0 g/cm^3$ while the density of solid water (ice) is $0.9 g/cm^3$. This allows ice to float on water and insulates the water below from the cold.
- High surface tension. Helps carry water from roots to leaves (capillary action).
- Takes a great deal of energy to vaporize. Sweat is a very effective cooling mechanism.
- Universal solvent. The polar nature of its bonds allows it to dissolve a wide variety of substances.



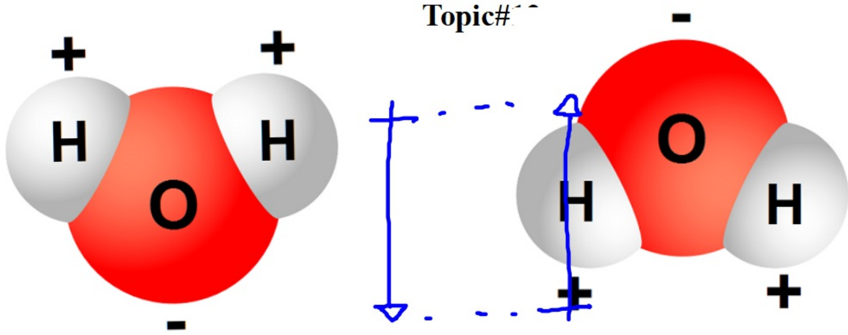
## Solid-Liquids
## Topic#12

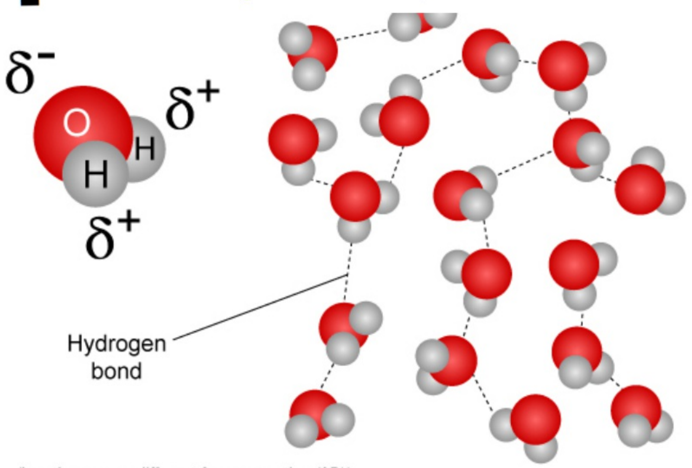
**Water**




Anchorage, AK is located in a temperate rain forest due to the warm ocean current in the Gulf of Alaska. This large, warm body of water moderates the temperature of the surrounding land mass making the average temperature around anchorage higher than many continental cities/regions.

**Water**



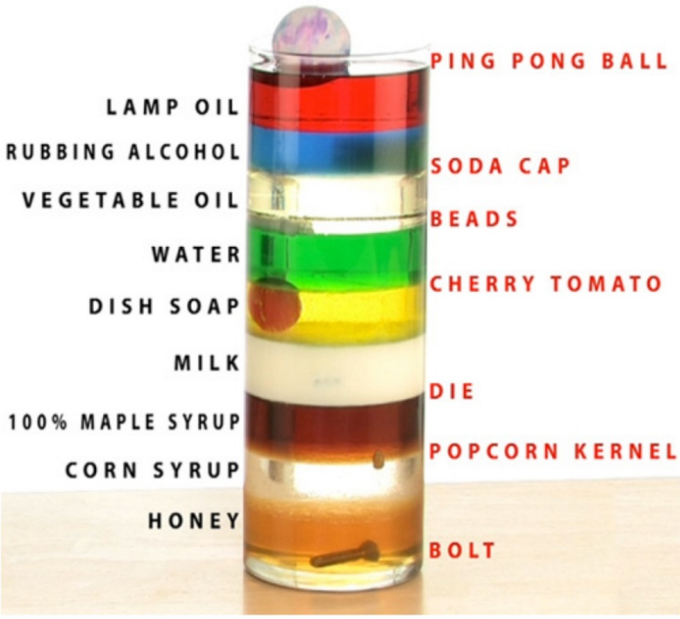

(length appears different for perspective (3D))

Dept. Biol. Penn State ©2002

## Solid-Liquids
## Topic#12

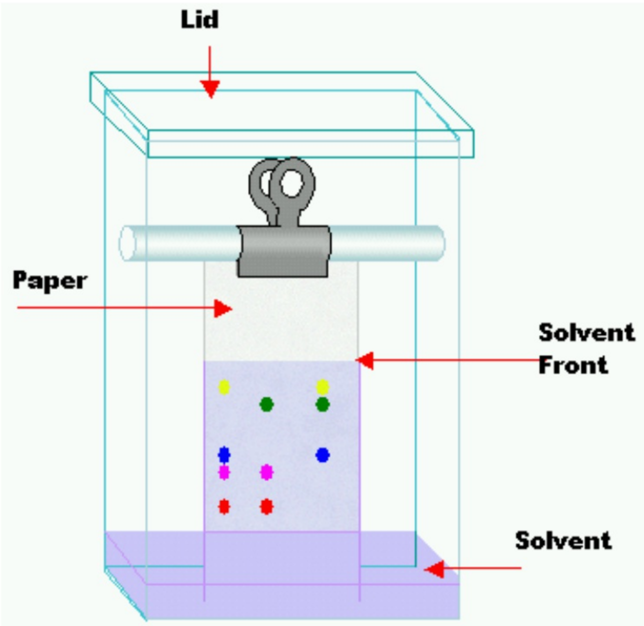
**Water**

Density Column




**Water**

Chromatography

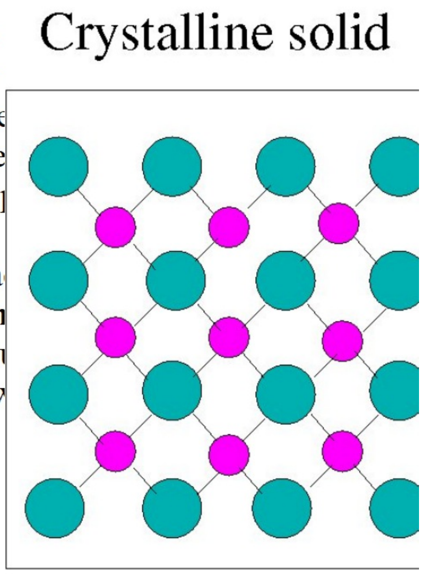


**Solid-Liquids**
**Topic#12**

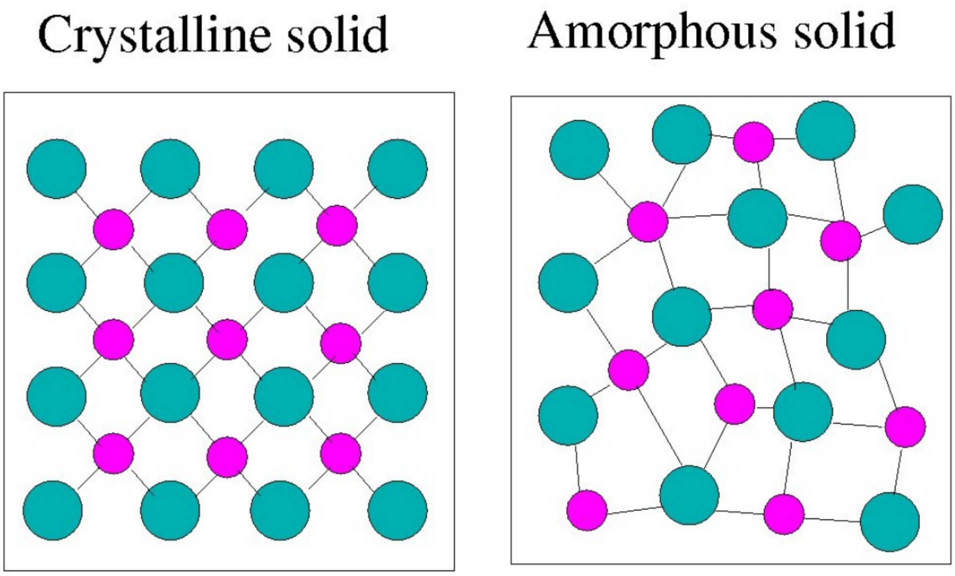
**Solids**

**Properties of Solids**

1) Low KE ($KE_{(s)} < KE_{(l)} < KE_{(g)}$)
2) Particles are fixed. Only movement is vibrational and rotational.
3) Definite shape and volume.
4) High density.
5) Held together with LDF, dipole-dipole, covalent, ionic, and metallic
   - Metallic/Network covalent > ionic > H-bonding > dipole-dipole >
   strongest (high mp) we
6) Definite **melting point** (mp) - transition from solid to liquid with the
   - Exceptions - amorphous solids which melt over a range of tem
7) Virtually incompressible
8) Low rate of diffusion - if it occurs, it only occurs at the surface of ea
9) Two general types of solids - **crystalline** (uniform structure) and **am**
   a) **Crystalline solids** - crystal structure (formed from repeating
   b) **Amorphous solids** - particles are arranged randomly ($SiO_2$, w
   liquids. Melt over a range of temperatures.



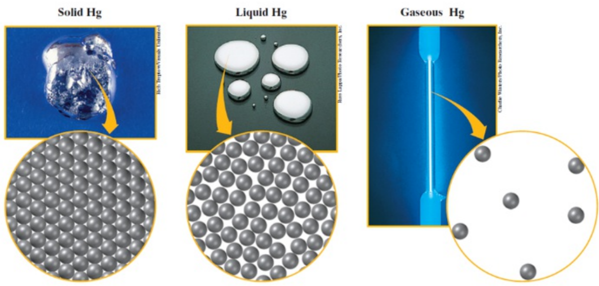
**Solids**



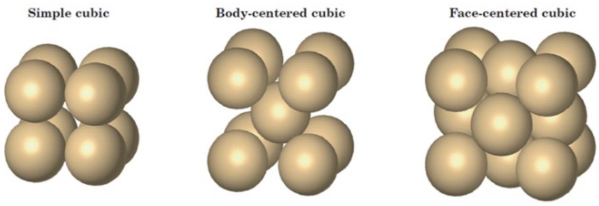
# Solid-Liquids
# Topic#12

## Solids

Prince Rupert Drop




### Unit Cells



## Solid-Liquids
## Topic#12

$\overset{2+}{Mg}\ \overset{2-}{O}$ $\overset{1+}{Na}\ \overset{1-}{F}$

**Solids**

**Crystalline Solids**

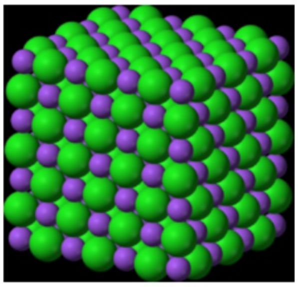
1) Ionic crystals - pattern of positive and negative ions arranged in a crystal lattice.
   - strong electrostatic forces between cations and anions
   - high mp, brittle, hard, conduct electricity only in molten form otherwise are good insulators (non-conductors)
2) Covalent Network Crystals - particles are covalently bonded to each other. Covelent bonding extends through the solid.
   - $C_{diamond}$, quartz, SiC (silicon carbide), etc.
   - Hard and brittle.
3) Metallic Crystals - metal cations surrounded by a sea of delocalized electrons.
   - Electrons come from metal atoms but belong to the crystal as a whole.
   - Free electrons allow for thermal and electrical conductivity.
4) Covalent Molecular Crystals - nonpolar/polar molecular solids held together by intermolecular forces.
   - nonpolar solids held together by LDFs ($He/H_2/CO_2/CH_4/C_6H_6$, etc.)
   - polar solids held together with dipole-dipole forces or hydrogen bonding.
   - much weaker than the covalent bonds between atoms in the molecules.
   - low mp/low bp/easily vaporized/soft/good insulators.

# Solid-Liquids
## Topic#12

**Solids**

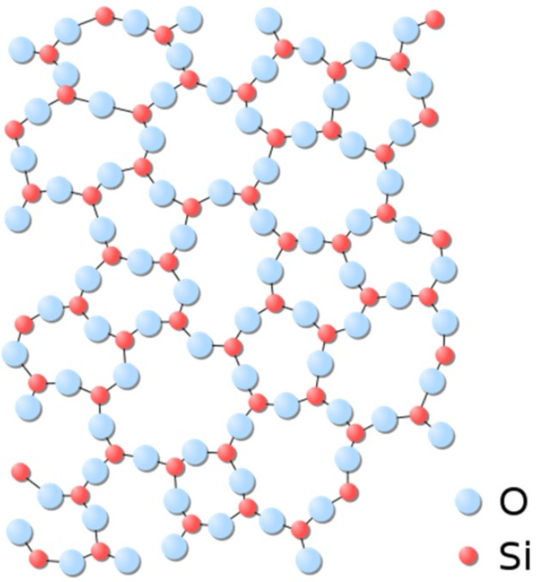
### Melting and Boiling Points of Representative Crystalline Solids

| Type of Substance | Formula | MP(°C) | BP(°C) (1 atm) |
|---|---|---|---|
| Ionic | NaCl | 801 | 1413 |
| | $MgCl_2$ | 1266 | 2239 |
| Covalent Network | $(SiO_2)_x$ | 1610 | 2230 |
| | $C_x$ (diamond) | 3500 | 3930 |
| Metallic | Hg | -39 | 357 |
| | Cu | 1083 | 2567 |
| | Fe | 1535 | 2750 |
| | W | 3410 | 5660 |
| Covalent Molecular (Polar) | $NH_3$ | -78 | -33 |
| | $H_2O$ | 0 | 100 |
| Covalent Molecular (Nonpolar) | $H_2$ | -259 | -253 |
| | $O_2$ | -218 | -183 |
| | $CH_4$ | -182 | -164 |
| | $CCl_4$ | -23 | 77 |
| | $C_6H_6$ | 6 | 80 |



**Solids**

**Amorphous Solids**

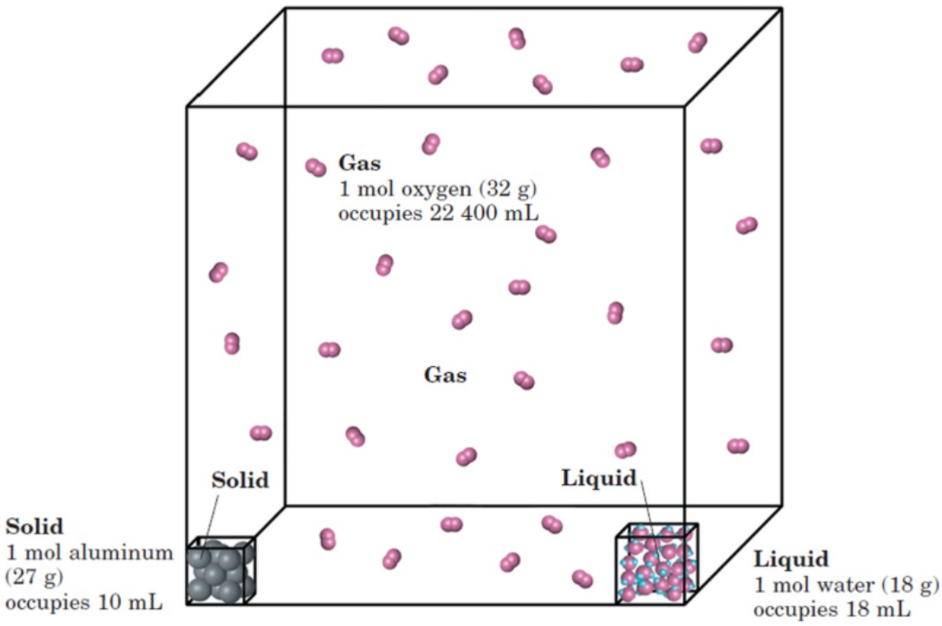



Note the disorganized structure intrinsic in an amorphous solid.

## Solid-Liquids
## Topic#12

**Solids**

### Comparing Relative Volumes for a Solid, Liquid, and Gas



**Solids**

**Comparing the Bonds between N-H, N-Cl, and N-I**

Bond Energies (kJ/mol)

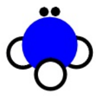
$BE_{N\text{-}I}$ (159) < $BE_{N\text{-}Cl}$ (200) < $BE_{N\text{-}H}$ (391)

Bond Lengths (pm)

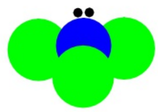
$BL_{N\text{-}I}$ (222) > $BL_{N\text{-}Cl}$ (191) > $BL_{N\text{-}H}$ (101)

The BL coupled with the small number of protons in the nucleus leads to the instability of the N-I bond. Even though $NI_3$ is a solid at romm temperature, it explodes on contact.

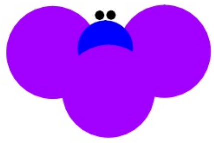
Video on Exploding $NI_3$
https://youtu.be/JRhvnlQHKc0



$NH_3$

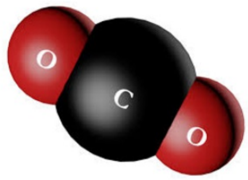


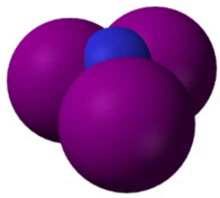
$NCl_3$



$NI_3$

**Liquified $CO_2$**

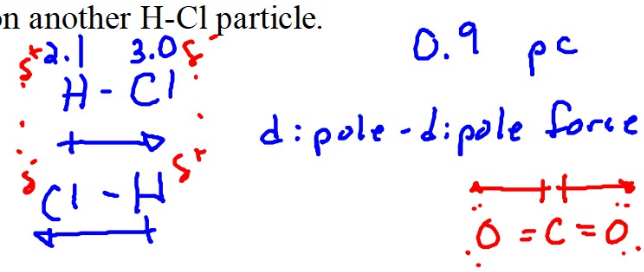


**Nitrogen Tiiodide
Contact Explosive**

## Intermolecular Forces

1) Holds liquids/solids to a specific volume.
2) Vary in strength from strong (metallic/network covalent) to weak (london dispersion forces (LDF)).
3) **Dipole forces -** polar molecules with a dipole
   - Remember a dipole is created in a polar covalent bond where the $0.4 < EN_{\text{diff}} < 1.7$.
   - For H-Cl molecule, a "force" of attraction between the partial positive H and the partial negative Cl on another H-Cl particle.

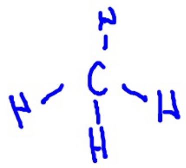


**Solids - Intermolecular Forces**

4) **Hydrogen bonding (dipole-dipole force) -** intermolecular force between the H on one particle and a highly electronegative element on another particle (N/O/F).
   - $H_2O$, HF, HCl, $NH_3$, etc.

5) **London dispersion forces (LDF) -** an intermolecular force between two particles involving the "induced dipole" effect.
   - **Induced dipole** - created when one electron cloud of a particle "pushes" the electron cloud of another particle briefly exposing the positive nucleus to the opposing elctron cloud. LDF forces are ubiquitous in all solids and liquids, but are the only force involved in the solidification/liquification of non-polar molecules and noble gases ($CO_2$/$CH_4$/$C_2H_6$/$CCl_4$/$SO_3$/He/Ar/Xe, etc). The larger the particle the stronger the LDFs.

- **Example:** As the halogens increase in atomic mass, how do their states at 25°C change?



npc

$CH_4$ 16.05 g/mol
$CCl_4$ 154
$CBr_4$ 332
$CI_4$ 520

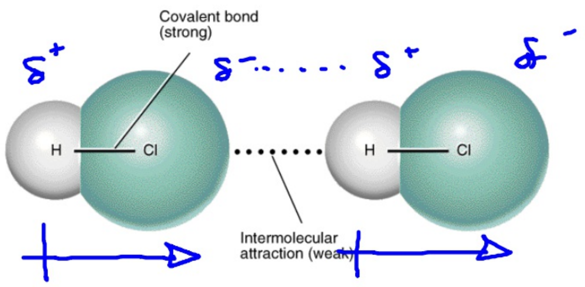
**Solids - Intermolecular Forces**

**Video on the sublimation and deposition of $I_2$.**

# Solid-Liquids
# Topic#12

## Solids

Dipole-Dipole force

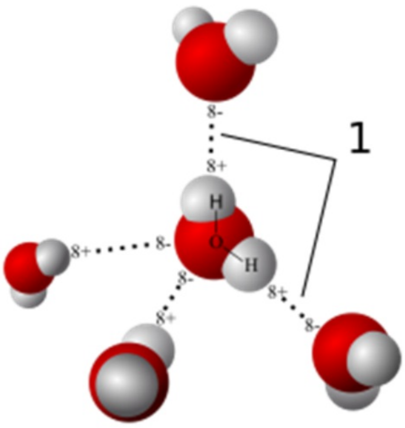


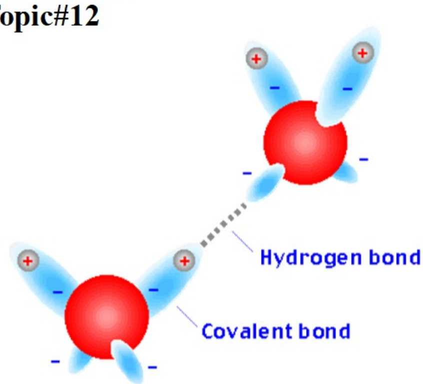

EN (Cl) = 3.0
EN (H) = -2.1
$EN_{diff}$ 0.9

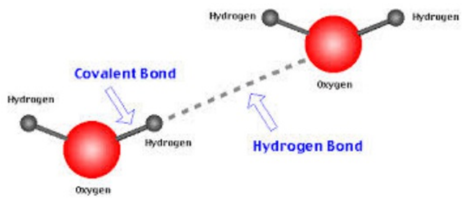
# Solid-Liquids
# Topic#12

## Solids

Hydrogen Bonding



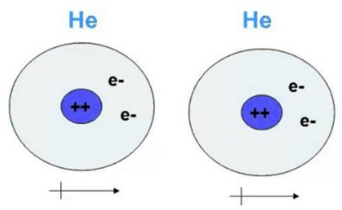



**Solids - Intermolecular Forces**

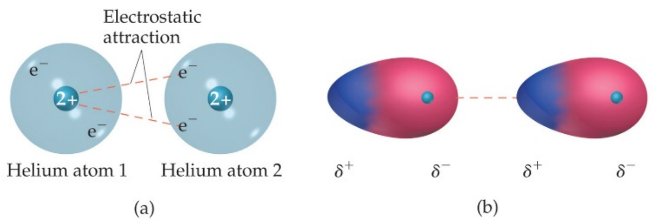
## London Dispersion Forces

Electron correlation leads to "instantaneous dipoles."

If helium atoms move slowly enough (at extremely low temperatures), this weak attractive force allows atoms to stick together (condense), leading to the gas-to-liquid phase transition.



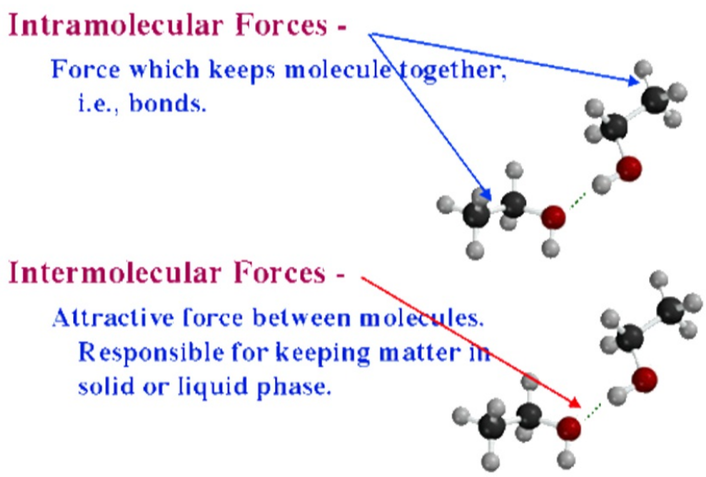

## Process for Induced Dipoles (LDFs)




Copyright © 2009 Pearson Prentice Hall, Inc.

**Solid-Liquids**
**Topic#12**

**Solids - Intermolecular Forces**




| Intramolecular | Particles |
|---|---|
| Ionic | $Na^+$ $Cl^-$ |
| Covalent | H-H |
| Metallic | Fe-Fe |

| Intermolecular | Particle Interactions |
|---|---|
| Ion-Dipole | $Na^+ \cdots OH_2$ |
| H-Bonding | $HOH \cdots OH_2$ |
| Dipole-Dipole | $HCl \cdots HCl$ |
| Dispersion | $H\text{-}H \cdots H\text{-}H$ |

**Solids - Intermolecular Forces**

### Comparison of the Energies Associated with Bonding (Intramolecular) Forces and Intermolecular Forces

| Force | Model | Basis of Attraction | Energy (kJ/mol) | Example |
|---|---|---|---|---|
| **Intramolecular** | | | | |
| Ionic | | Cation–anion | 400–4000 | NaCl |
| Covalent | | Nuclei–shared $e^-$ pair | 150–1100 | H—H |
| Metallic | | Cations–delocalized electrons | 75–1000 | Fe |
| **Intermolecular** | | | | |
| Ion-dipole | | Ion charge–dipole charge | 40–600 | $Na^+ \cdots OH_2$ |
| H bond | $\overset{\delta^-}{-A}-\overset{\delta^+}{H}\cdots\cdots:\overset{\delta^-}{B}-$ | Polar bond to H–dipole charge (high EN of N, O, F) | 10–40 | $:\ddot{O}(H)-H\cdots:\ddot{O}(H)-H$ |
| Dipole-dipole | | Dipole charges | 5–25 | I—Cl····I—Cl |
| Ion-induced dipole | | Ion charge–polarizable $e^-$ cloud | 3–15 | $Fe^{2+}\cdots O_2$ |
| Dipole-induced dipole | | Dipole charge–polarizable $e^-$ cloud | 2–10 | H—Cl····Cl—Cl |
| Dispersion (London) | | Polarizable $e^-$ clouds | 0.05–40 | F—F····F—F |

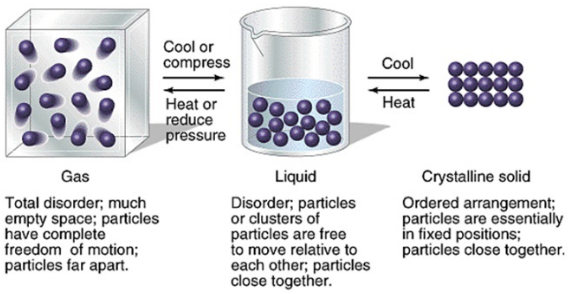
# Solid-Liquids
# Topic#12

## Phase Changes

**Phase** - any part of a system with uniform composition.
- solid (*s*), liquid (*l*), and gas (*g*)
- also called state of substance

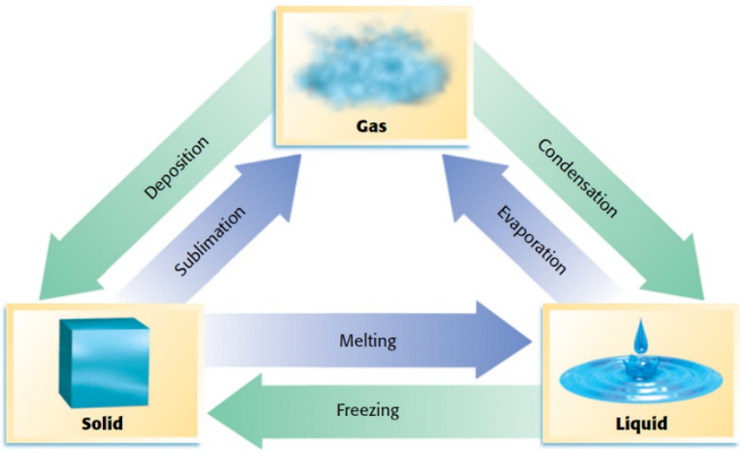
**Phase change** - a physical change from one phase to another.

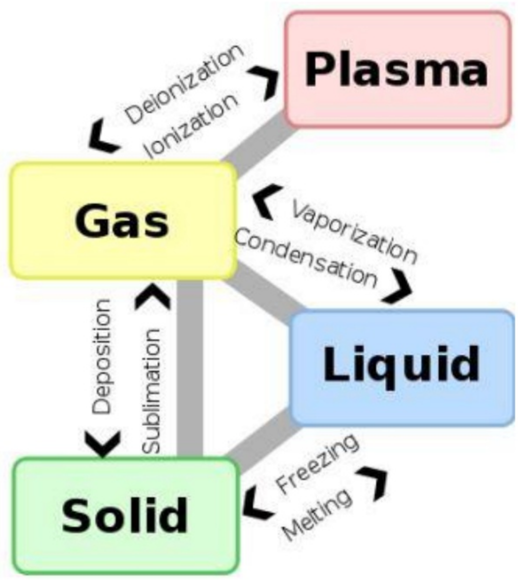
| | | |
|---|---|---|
| **melting** - solid to liquid | endothermic | $H_2O(s) \rightarrow H_2O(l)$ |
| **freezing** - liquid to solid | exothermic | $H_2O(l) \rightarrow H_2O(s)$ |
| **vaporization** - liquid to gas | endothermic | $Br_2(l) \rightarrow Br_2(g)$ |
| **condensation** - gas to liquid | exothermic | $H_2O(g) \rightarrow H_2O(l)$ |
| **sublimation** - solid to gas | endothermic | $CO_2(s) \rightarrow CO_2(g)$ |
| **deposition** - gas to solid | exothermic | $H_2O(g) \rightarrow H_2O(s)$ |
| **boiling** - liquid to gas (bubbles) @ 100°C and 1 atm | endothermic | $H_2O(l) \rightarrow H_2O(g)$ |
| **evaporation** - liquid to gas (below boiling point) | endothermic | $H_2O(l) \rightarrow H_2O(g)$ |



# Solid-Liquids
# Topic#12

## Solids



**Equilibrium Vapor Pressure**

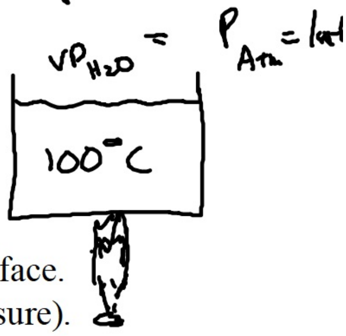
liquid + *E* <--> gas

## Equilibrium vapor pressure

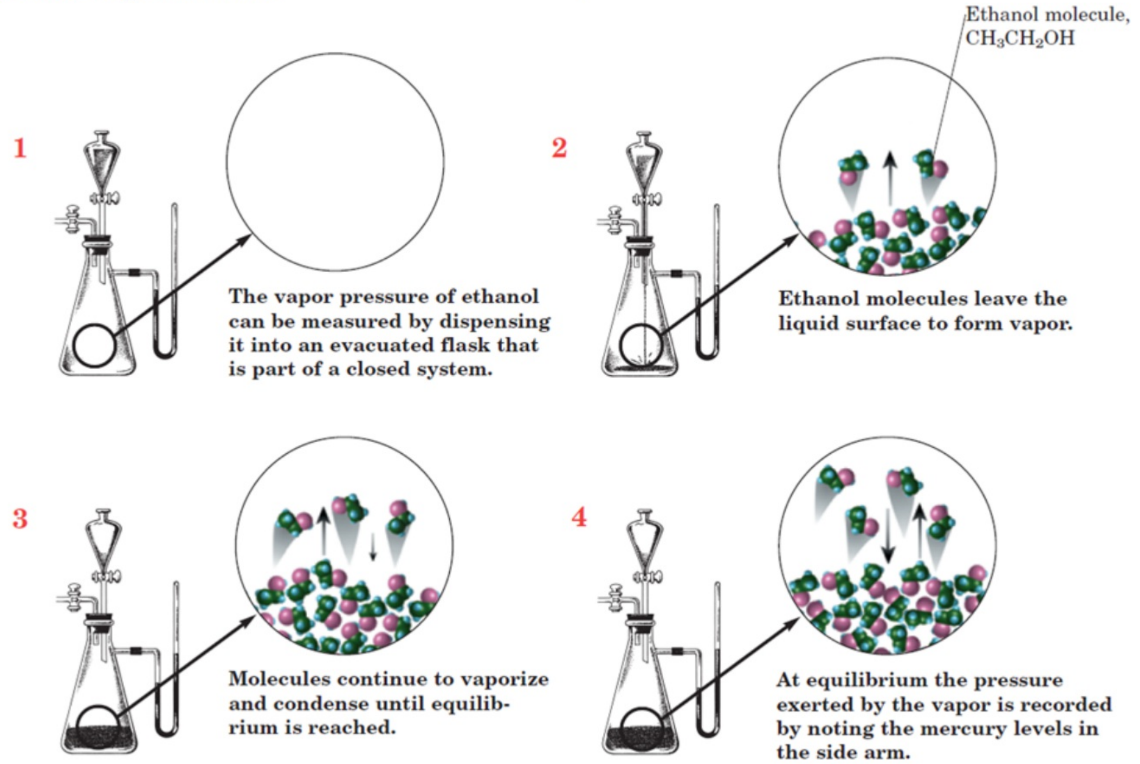
- closed system
- equilibrium between liquid and vapor above liquid
- the rate of vaporization (evaporation) and rate of condensation are equal
- equilibrium *VP* changes with temperature
- usually, the warmer the system the higher the *VP*.
- volatile liquids evaporate readily(high equil *VP*) (alcohol, water)
- nonvolatile liquids do not evaporate readily (low equil *VP*) (oil)


## Boiling

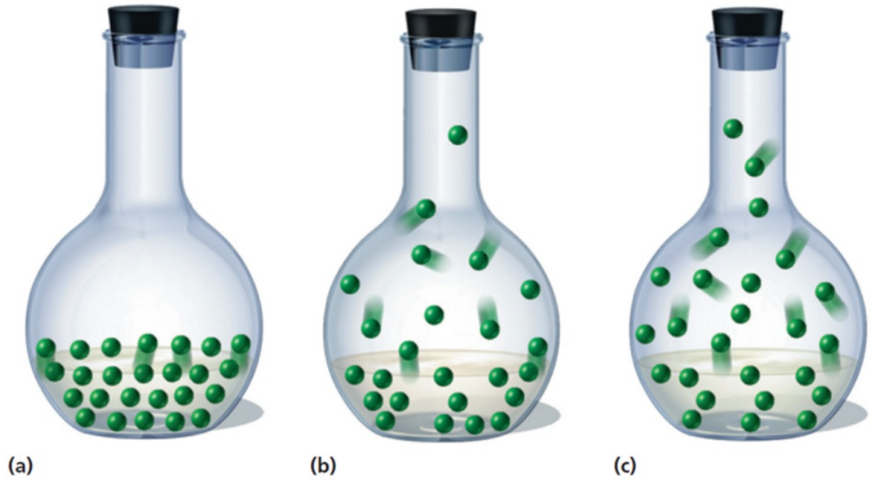
- normal *bp* occurs at 1atm.
- conversion of liquid to vapor within the liquid (bubbles) as well as at surface.
- point where equil *VP* is equal to the atmospheric pressure (ambient pressure).
- temperature of liquid DOES NOT change during the transition to vapor at the bp



## Equilibrium Vapor Pressure



## Equilibrium Vapor Pressure

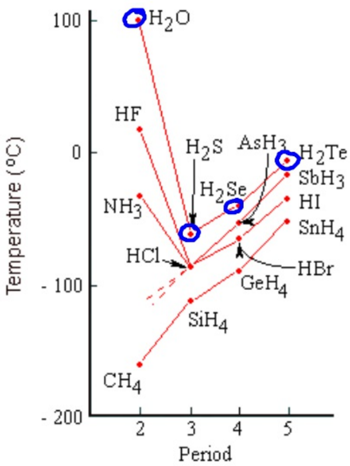



A liquid-vapor equilibrium develops in a closed system. (a) At first there is only liquid present, but molecules are beginning to evaporate. (b) Evaporation continues at a constant rate. Some vapor molecules are beginning to condense to liquid. (c) Equilibrium has been reached between the rate of condensation and the rate of evaporation.

## Liquids

Boiling Point

- The stronger the intermolecular forces the higher the b.p..
- Usually, the heavier the particle the higher the b.p..
  - Please order the b.p. of $I_2$, $Cl_2$, $F_2$, and $Br_2$ from highest to lowest.
- High EN elements bonded to H have unusually high b.p. for their group.




Electronegativities

| IV | V | VI | VII |
|---|---|---|---|
| C 2.5 | N 3.0 | O 3.5 | F 4.0 |
| Si 1.8 | P 2.1 | S 2.5 | Cl 3.0 |
| Ge 2.0 | As 2.2 | Se 2.6 | Br 2.8 |
| Sn 1.8 | Sb 2.0 | Te 2.1 | I 2.5 |
| Pb 2.3 | Bi 2.0 | Po - | |

$I_2 > Br_2 > Cl_2 > F_2$

2(53) 2(35) 2(17) 2(9)

LDF

O-H
$3.5 - 2.1 = 1.4$

S-H
$2.5 - 2.1 = 0.4$

Se-H
$2.6 - 2.1 = 0.5$

**Liquids**

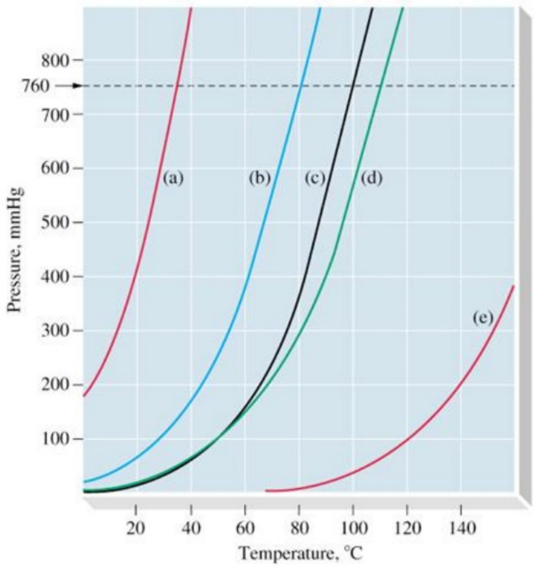
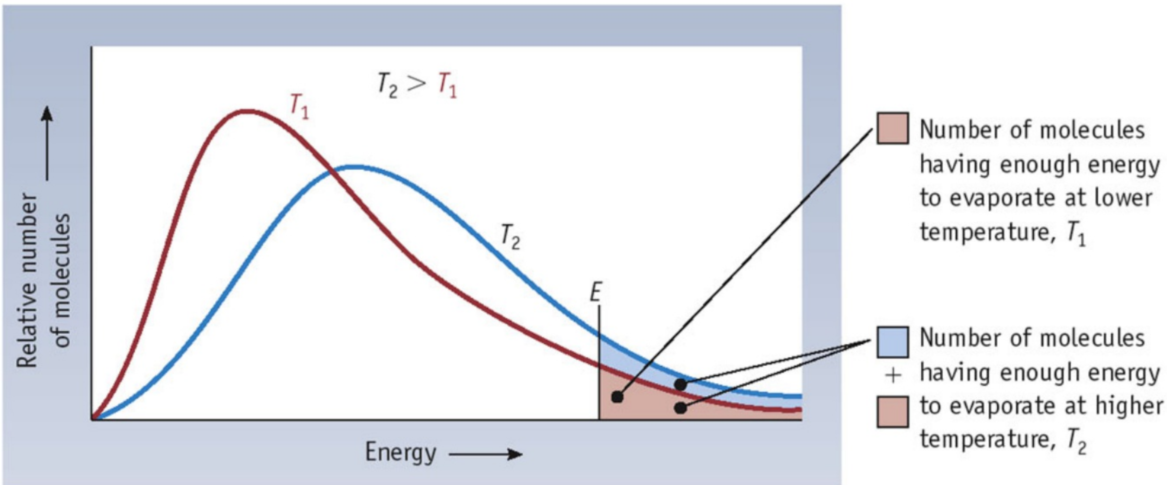



**Figure 1. Typical vapour pressure-temperature curves (a) Diethyl ether, (b) Benzene, (c) Water, (d) Toluene, (e) Aniline. 760 mmHg is the standard atmospheric pressure.**

**Liquids**




© 2006 Brooks/Cole - Thomson

**Energy Distribution of Molecules in a Liquid at Different Temperatures ($T_1$ and $T_2$)**

## Phase Diagrams

### Phase Diagram

A graph of $P$ vs $T$ showing the conditions under which phases of a substance exist.

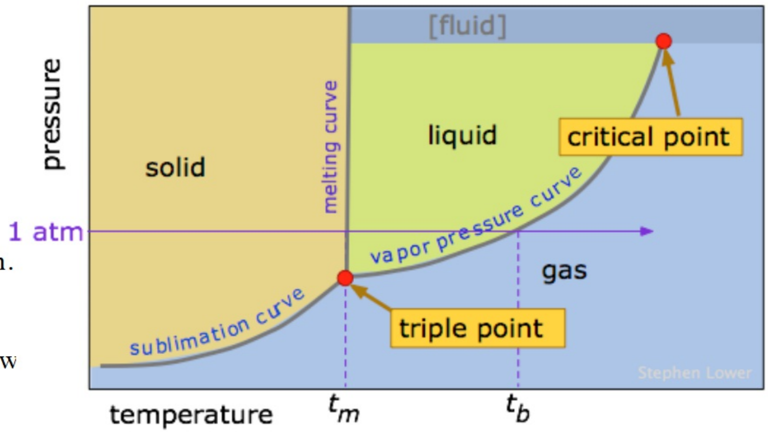
**Phases** - $s/l/g$ (clockwise on graph)

**Triple point** - $T$ and $P$ where a substances solid, liquid, and gas state are in equilibrium.

**Critical point** - critical $T$ and $P$ of a substance. At this point, a liquid can no longer exist. Too much KE in system to allow for intermolecular bonds to exist.

**Critical Temperature** - maximum temperature the liquid state of a substance can exist.

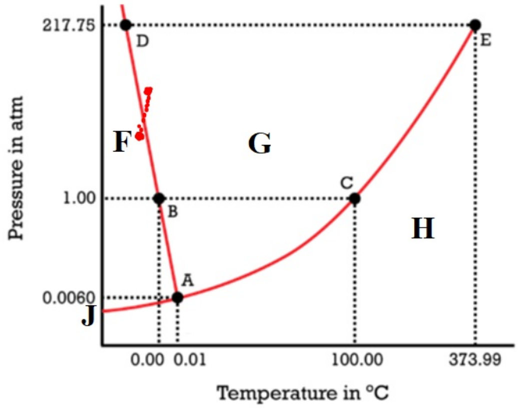
**Critial Pressure** - the lowest pressure that can be placed on a liquid at critical $T$.



## Solid-Liquids
## Topic#12

## Phase Diagrams

### Topic#12 Solds-Liquids SW#1 - Phase Diagram




1. After studying the phase diagram to the right, identify the substance.

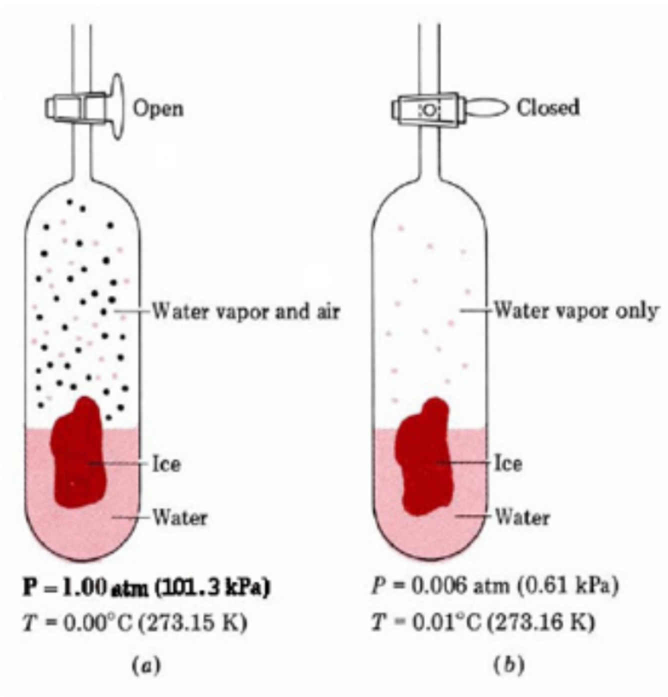
___________________

2. What letter represents the solid phase? _________
3. What letter represents the liquid phase? _________
4. What letter represents the gas phase? _________
5. What letter represents the substance's normal freezing point? _________
6. What is its normal freezing point? _________
7. What letter represents the substance's normal boiling point? _________
8. What is its normal boiling point? _________
9. What is meant by "normal" freezing or boiling point?

_________________________________

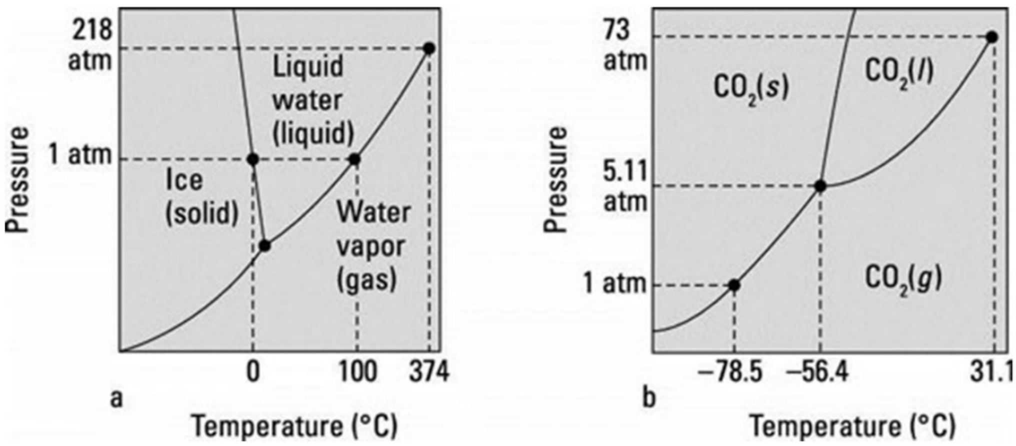
10. What letter represents the substances triple point? _________
11. Identify the conditions the substance requires to be at its triple point _________ & _________
12. What letter represents the substances critical point? _________
13. Identify the conditions the substance requires to be at its critical point _________ & _________
14. Which phase of the substance is the densest? (solid / liquid / gas)
15. What two phase changes occur along the line between A and D? ________________ & ________________
16. What is the name of this line? ____________________________________________
17. What two phase changes occur along the line between A and E? ________________ & ________________
18. What is the name of this curve? ____________________________________________
19. What two phase changes occur along the line between A and J? ________________ & ________________
20. What is the name of this curve? ____________________________________________

**Triple Point**



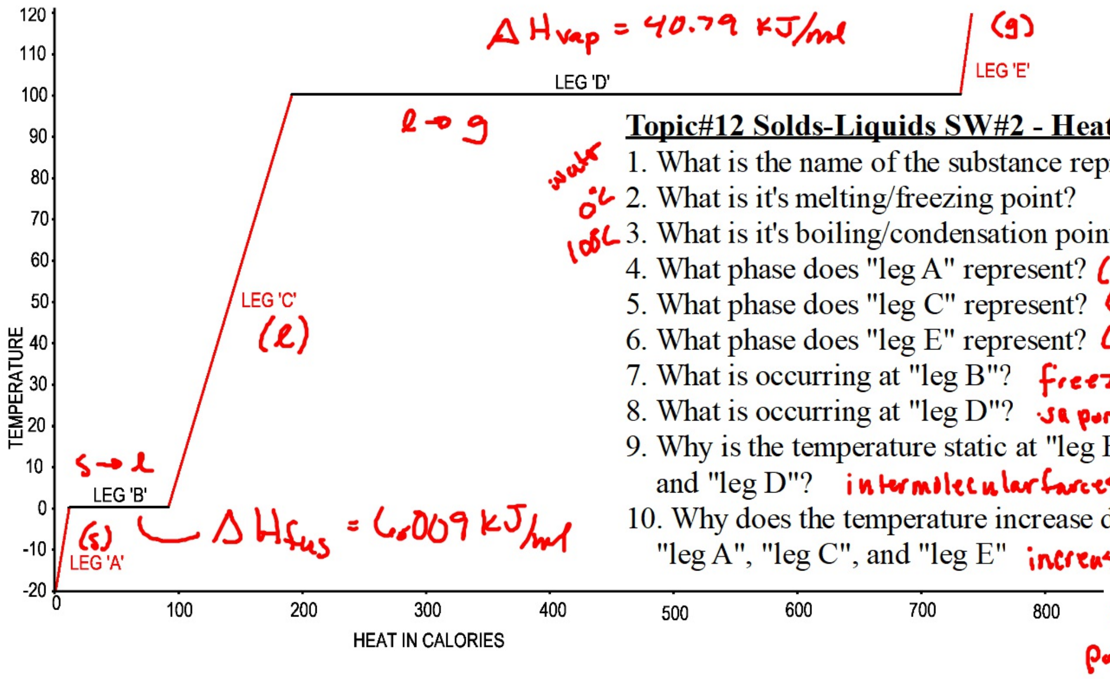
**Phase Diagrams**

## Phase Diagram
## $H_2O$ vs. $CO_2$



# Solid-Liquids
# Topic#12

## Heating Curve




**Topic#12 Solds-Liquids SW#2 - Heating Curve**

1. What is the name of the substance represented by the — water
2. What is it's melting/freezing point? — 0°C
3. What is it's boiling/condensation point? — 100C
4. What phase does "leg A" represent? (s)
5. What phase does "leg C" represent? (l)
6. What phase does "leg E" represent? (g)
7. What is occurring at "leg B"? freezing/melting
8. What is occurring at "leg D"? vaporization/condensation
9. Why is the temperature static at "leg B" and "leg D"? intermolecular forces
10. Why does the temperature increase during "leg A", "leg C", and "leg E" increasing the KE of the particles in a phase

**Phase Change Energy**

**solid + energy <--> liquid**

## Freezing/Melting

- normal freezing/melting point - temperature at which the solid and liquid are in equilibrium at 1atm (760torr/101.3kPa) of pressure.

**Molar Heat of Fusion ($\Delta H_f$)**

- the energy, as heat, needed to melt 1mol of a solid at the solid's normal mp.
- water's molar heat of fusion is 6.009kJ/mol

enthalpy ($H$)

**Molar Heat of Fusion**

| | | |
|---|---|---|
| $H_2O(s)$ + 6.009kJ/mol --> $H_2O(l)$ | $+\Delta H_f$ | endothermic |
| $H_2O(l)$ --> $H_2O(s)$ + 6.009kJ/mol | $-\Delta H_f$ | exothermic |

**Phase Change Energy**

**liquid + energy <--> gas**

## Vaporization and Condensation

- normal boiling/condensation point - temperature at which the liquid and gas are in equilibrium at 1atm (760torr/101.3kPa) of pressure.

**Molar Enthalpy of Vaporization ($\Delta H_v$)**

- the energy needed to vaporize 1 mole of a substance at the bp under constant pressure
  - water's molar enthalpy of vaporization is 40.79kJ/mol

6.009 = ΔH_f

**Molar Heat of Vaporization**

$H_2O(l)$ + 40.79kJ/mol --> $H_2O(g)$ $+\Delta H_v$

$H_2O(g)$ --> $H_2O(l)$ + 40.79kJ/mol $-\Delta H_v$

**Solid-Liquids**
**Topic#12**

**Phase Change Energy**

**1 cal = 4.184J**

**1kcal = 1000cal**

**1Cal = 1000cal**

**Calorie** - the amount of heat needed to raise 1g of water 1°C

**Joule** - the amount of work done by a force of 1N moving through a distance of 1 meter. SI unit for energy.

**Energy Conversions Practice Problems**

11. How many joules are in 216cal? How many calories in 216J? How many Calories are in 34,745J? (Ans: 903J, 51.6cal, and ~~8.3043~~Cal)

Gvn
216 cal
216 Joules
34,745 Joules

NTK
4.184 J = 1 cal
1000 cal = 1 Calorie (Kcal)

Unk
J = ______ J
cal = ______ cal
Cal = ______ Cal

Solve:

$216\ cal \times \frac{4.184\ J}{1\ cal} = \boxed{903\ J}$

$216\ J \times \frac{1\ cal}{4.184\ J} = \boxed{51.6\ cal}$

$34{,}745\ J \times \frac{1\ cal}{4.184\ J} \times \frac{1\ Cal}{1000\ cal} = \boxed{8.3043\ Cal}$

**Solid-Liquids**
**Topic#12**

**Phase Change Energy**

$$q = \text{mol} \times \Delta H$$

$\Delta H_{fus} = 6.009 \text{ kJ/mol}$

## Molar Heat of Fusion Practice Problem

$1 \text{ m } H_2O = 18.02 \text{ g/mol}$

12. 23.15kJ of heat energy is needed to melt how many grams of water? Ans: 69.42g

Solve: $\frac{23.15 \text{ kJ}}{} \times \frac{1 \text{ mol}}{6.009 \text{ kJ}} \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 69.42 \text{ g } H_2O$

13. How much energy is needed to melt 8.91g of water? Ans: 2.97kJ

Solve: $\frac{8.91 \text{ g}}{} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{6.009 \text{ kJ}}{1 \text{ mol}} = 2.97 \text{ kJ}$

## Molar Heat of Vaporization Practice Problems

$\Delta H_{vap} = 40.79 \text{ kJ/mol}$

14. How much energy is needed to boil 623g of water? Ans: 1410kJ

$\frac{623 \text{ g}}{} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{40.79 \text{ kJ}}{\text{mol}} = 1410 \text{ kJ}$

15.A certain amount of water is boiled by the addition of 999kJ. Calculate this mass of water. (Ans: 441g)

$\frac{999 \text{ kJ}}{} \times \frac{\text{mol}}{40.79 \text{ kJ}} \times \frac{18.02 \text{ g}}{\text{mol}} = 441 \text{ g}$

## Solid-Liquids
## Topic#12

**Specific Heat**

**Specific Heat ($S$) - the energy needed to raise 1 gram of a substance by 1°C.**

$$q = m \times S \times \Delta T = J$$

$$\Delta T = T_f - T_i$$

**$q$ is the heat gained/released by the system**

**m is the mass of the substance**

**$S$ is the specific heat of the substance (varies by state)**

**$\Delta T$ is the temperature change of the system**

**List of a few substances specific heats.**

- $S_{H2O}$ (l) = 4.184 J/(g-°C)
- $S_{H2O}$ (s) = 2.03
- $S_{H2O}$ (g) = 1.87
- $S_{Fe}$ = 0.449
- $S_{Cu}$ = 0.385
- $S_{Hg}$ = 0.129

**Specific Heat Practice Problem**

16. How much energy is needed to raise 34g of water from 25°C to 98°C? (Ans: $1.0 \times 10^4$J)

Given
34g = m
$S_{H_2O(l)} = 4.184 \frac{J}{g \cdot °C}$
$\Delta T = T_f - T_i = 98 - 25 = 73°C$

NTK
$q = m \times S \times \Delta T$

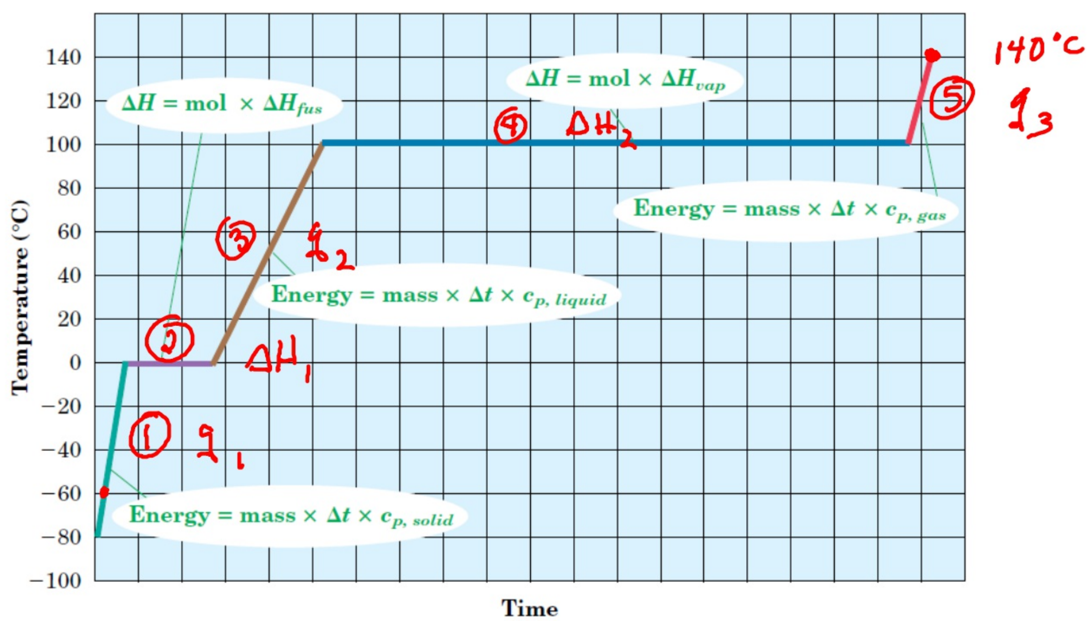
Unk
$q = \underline{1.0 \times 10^4}$ J

Solve:
$q = (34)(4.184)(73) = 1.0 \times 10^4$ J
$g \times \frac{J}{g \cdot °C} \times °C = J$

**Heating Curve Calculations**

## Calculating Energy Changes for a Heating Curve

$q_1 + \Delta H_1 + q_2 + \Delta H_2 + q_3 = q_T$

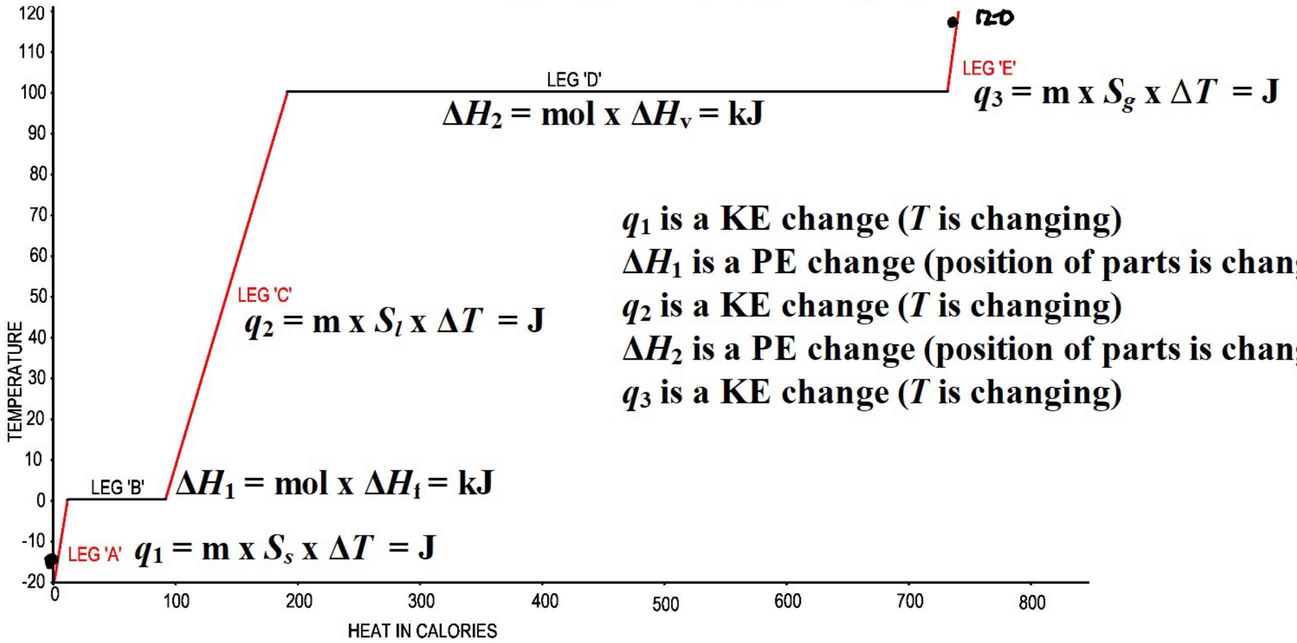


**Solid-Liquids**
**Topic#12**

## Heating Curve Calculations

Method used to calculate the energy needed to heat a substance in the solid state to its gaseous state.

$$q_T = q_1 + \Delta H_1 + q_2 + \Delta H_2 + q_3$$




**$q_1$ is a KE change ($T$ is changing)**
**$\Delta H_1$ is a PE change (position of parts is changing)**
**$q_2$ is a KE change ($T$ is changing)**
**$\Delta H_2$ is a PE change (position of parts is changing)**
**$q_3$ is a KE change ($T$ is changing)**

## Solid-Liquids
## Topic#12

### Heating Curve Calculations

$S_{H2O}(l) = 4.184\ J/^oC\text{-}g$

$S_{H2O}(g) = 1.996 J/^oC\text{-}g$

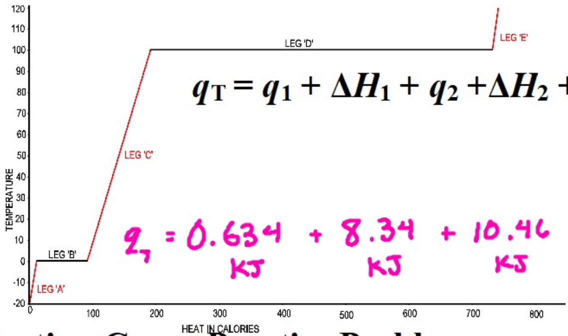
$S_{H2O}(s) = 2.108\ J/^oC\text{-}g$

$\Delta H_f = 6.009 kJ/mol$

$\Delta H_v = 40.79 kJ/mol$

Convert:

$25.00g \times \frac{1\,mol}{18.02g} = 1.387\ mol$




$q_T = q_1 + \Delta H_1 + q_2 + \Delta H_2 + q_3$

$\Delta H_1 = \text{mol} \times \Delta H_{fus}$

$\Delta H_2 = \text{mol} \times \Delta H_{vap}$

$q = m \times S \times \Delta T$

$\Delta T = T_f - T_i$

$q_T = 0.634\ kJ + 8.34\ kJ + 10.46\ kJ + 56.58\ kJ + 1.497\ kJ$

$= 77.5\ kJ$

### Calculations Involving a Heating Curve Practice Problem

17. Calculate the amount of energy required to heat up 25.00g of ice from -12.0°C to 130.0°C.

$\Delta T_s = 0 - (-12) = 12°C$ (ice) $\quad \Delta T_g = 130 - 100 = 30°C \quad q_T = q_1 + \Delta H_1 + q_2 + \Delta H_2 + q_3$

$\Delta T_l = 100 - 0 = 100°C$

$q_1 = m \times S_s \times \Delta T = J = (25g) \times (2.108) \times (12) = 632.4\ J = 0.634\ kJ$

$\Delta H_1 = \text{mol} \times \Delta H_f = kJ \quad 1.387\,mol \times 6.009\ kJ = 8.34\ kJ$

$q_2 = m \times S_l \times \Delta T = J = (25g) \times (4.184_{(l)}) \times (100) = 10{,}460\ J = 10.46\ kJ$

$\Delta H_2 = \text{mol} \times \Delta H_v = kJ \quad 1.387\,mol \times 40.79\ kJ = 56.58\ kJ$

$q_3 = m \times S_g \times \Delta T = J = (25g) \times (1.996_{(gas)}) \times (30) = 1497\ J = 1.497\ kJ$

## Solid-Liquids
## Topic#12

### Heating Curve Calculations

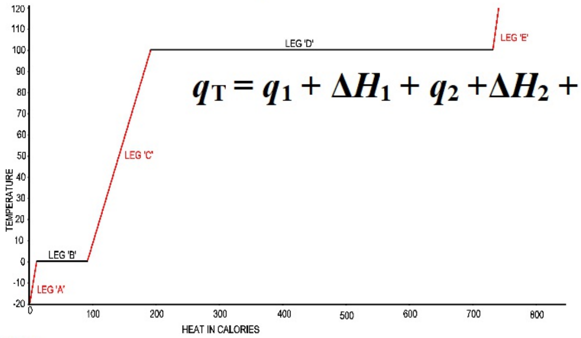
**$S_{H2O}(l)$ = 4.184 J/°C-g**

**$S_{H2O}(g)$ = 1.996J/°C-g**

**$S_{H2O}(s)$ = 2.108 J/°C-g**

**$\Delta H_f$ = 6.009kJ/mol**

**$\Delta H_v$ = 40.79kJ/mol**




$$q_T = q_1 + \Delta H_1 + q_2 + \Delta H_2 + q_3$$

$$\Delta H_1 = \text{mol x } \Delta H_{fus}$$

$$\Delta H_2 = \text{mol x } \Delta H_{vap}$$

$$q = \text{m x S x } \Delta T$$

$$\Delta T = T_f - T_i$$

### Starter - Energy Conversions

1. How much energy is released when 110.grams of copper ($S_{Cu}$ = 0.385J/°C-g) is cooled from 350.°C to 152°C? (Ans: 8390J)
2. 784.5g of ice is melted. How much energy did this require? (Ans: 261.6kJ)
3. The amount of water in question #2 is vaporized. Calculate the energy needed to convert the liquid to gas. (Ans: 1776kJ)
4. Calculate the amount of energy required to heat up 15.00g of ice from -22.0°C to 120.0°C. (Ans: 46.4kJ)

## Solid-Liquids
## Topic#12

**Heating Curve Calculations**

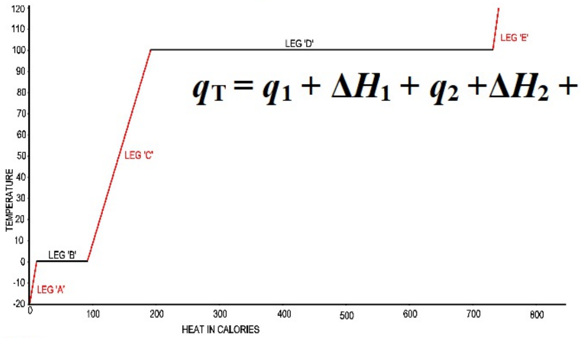
**$S_{H2O}(l)$ = 4.184 J/°C-g**

**$S_{H2O}(g)$ = 1.996J/°C-g**

**$S_{H2O}(s)$ = 2.108 J/°C-g**

**$\Delta H_f$ = 6.009kJ/mol**

**$\Delta H_v$ = 40.79kJ/mol**




$$q_T = q_1 + \Delta H_1 + q_2 + \Delta H_2 + q_3$$

$$\Delta H_1 = \text{mol x } \Delta H_{fus}$$

$$\Delta H_2 = \text{mol x } \Delta H_{vap}$$

$$q = \text{m x S x } \Delta T$$

$$\Delta T = T_f - T_i$$

**Starter - Energy Conversions**

1. How much energy is released when 110.grams of copper ($S_{Cu}$ = 0.385J/°C-g) is cooled from 350.°C to 152°C? (Ans: 8390J)
2. 784.5g of ice is melted. How much energy did this require? (Ans: 261.6kJ)
3. The amount of water in question #2 is vaporized. Calculate the energy needed to convert the liquid to gas. (Ans: 1776kJ)
4. Calculate the amount of energy required to heat up 15.00g of ice from -22.0°C to 120.0°C. (Ans: 46.4kJ)

**Heating Curve Calculations**

**WS#7 Answer Key Phase Energy**

1. 0.105cal
2. $1.99x10^5$J
3. $2.00x10^5$cal
4. 0.0600kcal
5. $1.25x10^5$ cal
6. $3.35x10^3$J
7. $-4.71x10^3$J
8. $4.44x10^3$J
9.