

Unit 6: Thermochemistry, States of Matter and Intermolecular Forces

Chapter 10: Causes of Change

10.1: Energy Transfer

Energy (E): - the ability to do **work** or produce **heat**.

Thermodynamics: - the study of the inter action of heat and other kinds of energy.

Different Types of Energy:

1. **Radiant Energy**: - solar energy from the sun.
2. **Thermal Energy**: - energy associated with the random motion of atoms and molecules.
3. **Chemical Energy**: - sometimes refer to as **Chemical Potential Energy**. It is the energy stored in the chemical bonds, and release during chemical change.
4. **Potential Energy**: - energy of an object due to its position.

Heat (q): - the transfer of energy between two objects (internal versus surroundings) due to the difference in temperature.

Work (w): - when force is applied over a displacement in the same direction or a change in volume under the same pressure ($w = F \times d = -P \Delta V$).
- work performed can be equated to energy if no heat is produced ($E = w$). This is known as the Work Energy Theorem.

Internal Energy (E): - total energy from work and heat within a system.

$\Delta E = q + w$ <p>$\Delta E = \text{Change in System's Internal Energy}$ $q = \text{heat}$ $w = \text{work}$</p>

Temperature: - the average kinetic energy of all the particles in a substance.

- Temperature is NOT Heat. A massive substance with a low temperature can have a lot of internal heat. This is because there are a lot of particles and even though their kinetic energy is low, their TOTAL energy is large.

Enthalpy (H): - the amount of internal energy at a specific pressure and volume (when there is no work done).

$\Delta E = q + w \quad (\text{if } w = 0, \text{ then } \Delta E = q = H)$

Chemistry Unit 6: Thermochemistry, States of Matter and Intermolecular Forces

Physical Change: - when the temperature of a substance is raised or lower, no new substance is formed.

The only heat change is the **physical kinetic energy** of the substance as its molecules are moving or vibrating faster or slower.

- when a substance is changing phase, no new substance is formed. The only heat change is the **physical potential energy** of the substance as its intermolecular forces are being loosen or completely broken.

Heat Units: - the measuring units to measure heat or energy.

a. **Joules:** - the metric unit to measure heat or energy named after English physicist James Prescott Joule.

b. **Calories:** - the old imperial unit to measure heat or energy. (1 cal = 4.184 J)

Specific Heat (c_p): - the amount of heat needed to raise **one gram of substance** by one degree Celsius or one Kelvin.

- the higher the specific heat, the more each gram of substance can “hold” the heat.

- units are in J/(g • °C) or kJ/(kg • °C) ; J/(g • K) or kJ/(kg • K)

Molar Heat Capacity (C): - the amount of heat needed to raise **one mole of substance** by one degree Celsius or one Kelvin.

- the higher the molar heat capacity, the more heat each mole or each particle of a substance can “hold”.

- units are in J/(mol • °C) or kJ/(kmol • °C) ; J/(mol • K) or kJ/(kmol • K)

Physical Kinetic Change

$$q = mc_p\Delta T$$

$$q = nC\Delta T$$

q = Change in Heat (J or kJ)

ΔT = Change in Temperature (in °C or K)

m = mass (g or kg)

c_p = Specific Heat [J/(g • °C or K) or kJ/(kg • °C or K)]

n = moles (mol or kmol)

C = Molar Heat Capacity [J/(mol • °C or K) or kJ/(kmol • °C or K)]

Specific Heat and Molar Heat Capacity of Some Common Substances

Substance	Specific Heat J/(g • °C or K) or kJ/(kg • °C or K)	Molar Heat Capacity J/(mol • °C or K) or kJ/(kmol • °C or K)
Ice H ₂ O _(s)	2.01	37.4
Water H ₂ O _(l)	4.18	75.3
Steam H ₂ O _(g)	2.00	36.8
Ammonia NH _{3(g)}	2.06	35.1
Methanol CH ₃ OH _(l)	2.53	81.1
Ethanol C ₂ H ₅ OH _(l)	2.44	112.4
Aluminum Al _(s)	0.897	24.2
Carbon (graphite) C _(s)	0.709	8.51
Copper Cu _(s)	0.385	24.5
Iron Fe _(s)	0.449	25.1
Silver Ag _(s)	0.235	46.3
Gold Au _(s)	0.129	25.3
Aluminum Chloride, AlCl _{3(s)}	0.690	92.0**

** Note that AlCl_{3(s)} has a molar specific capacity four times as large as Al_(s) because AlCl₃ has four times as many particles per mole as Al. (2 AlCl₃ → 2 Al + 6 Cl₂ 2 : 8 reactant to product particle ratio)

Relationship between Molar Heat Capacity and Specific Heat

(Molar Mass) × (Specific Heat) = (Molar Heat Capacity)

$$\left(\frac{\text{g}}{\text{mol}}\right) \times \left(\frac{\text{J}}{\text{g} \cdot \text{K}}\right) = \left(\frac{\text{J}}{\text{mol} \cdot \text{K}}\right)$$

Example 1: How much energy in kJ, is needed to heat 25.0 mol of water from 30.0°C to 75.0°C?

$C_{\text{water}} = 75.3 \text{ J}/(\text{mol} \cdot ^\circ\text{C})$

$n = 25.0 \text{ mol H}_2\text{O}$

$\Delta T = 75.0^\circ\text{C} - 30.0^\circ\text{C} = 45.0^\circ\text{C}$

$q = ?$

$q = nC\Delta T$

$q = (25.0 \text{ mol})(75.3 \text{ J}/(\text{mol} \cdot ^\circ\text{C}))(45.0^\circ\text{C}) = 84712.5 \text{ J}$

$q = 84.7 \text{ kJ}$

Example 2: When 1.80 mol of an unknown element cooled from 350 K to 270 K, it gave off 3.64 kJ of heat. Identify this unknown element by determining its molar heat capacity.

$q = 3.64 \text{ kJ} = 3640 \text{ J}$

$n = 1.80 \text{ mol}$

$\Delta T = 350 \text{ K} - 270 \text{ K} = 80 \text{ K}$

$C = ?$

$q = nC\Delta T$

$\frac{q}{n\Delta T} = C$

$C = \frac{(3640 \text{ J})}{(1.80 \text{ mol})(80 \text{ K})}$

$C = 25.3 \text{ J}/(\text{mol} \cdot \text{K})$

From the reference table above, the unknown substance is gold.

Example 3: Determine its final temperature of 0.157 mol of copper wire at an initial temperature of 15.0°C is heated using 2.27 kJ of energy.

$q = 2.27 \text{ kJ} = 2270 \text{ J}$

$n = 0.157 \text{ mol Cu}$

$C_{\text{Cu}} = 24.5 \text{ J}/(\text{mol} \cdot ^\circ\text{C})$

$T_i = 15.0^\circ\text{C}$

$\Delta T = T_f - 15.0^\circ\text{C} = ?$

$q = nC\Delta T$

$\frac{q}{nC} = \Delta T$

$\Delta T = \frac{(2270 \text{ J})}{(0.157 \text{ mol})(24.5 \text{ J}/(\text{mol} \cdot ^\circ\text{C}))}$

$\Delta T = 590^\circ\text{C}$

$\Delta T = T_f - 15.0^\circ\text{C}$

$590^\circ\text{C} = T_f - 15.0^\circ\text{C}$

$T_f = 590^\circ\text{C} + 15.0^\circ\text{C}$

$T_f = 605^\circ\text{C}$

Assignment

10.1 pg. 342 #1 to 4 (Practice); pg. 344 # 1 to 19

10.2: Using Enthalpy

System: - a part of the entire universe as defined by the problem.

Surrounding: - the part of the universe outside the defined system.

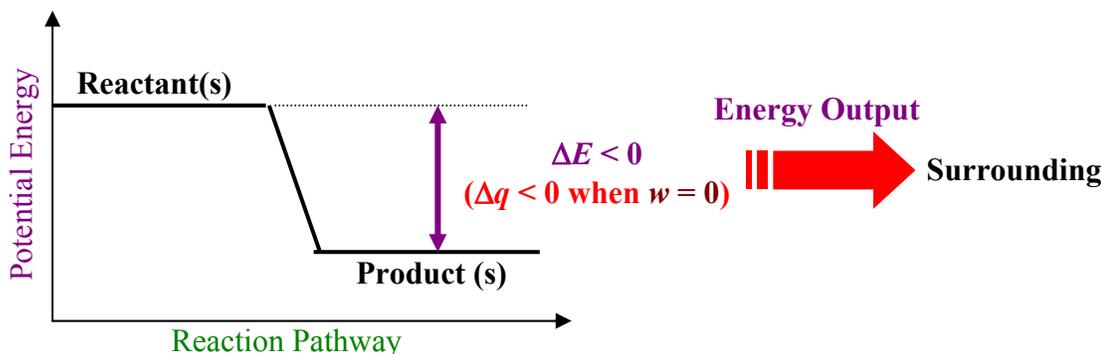
Open System: - a system where mass and energy can interchange freely with its surrounding.

Closed System: - a system where only energy can interchange freely with its surrounding but mass not allowed to enter or escaped the system.

Isolated System: - a system mass and energy cannot interchange freely with its surrounding.

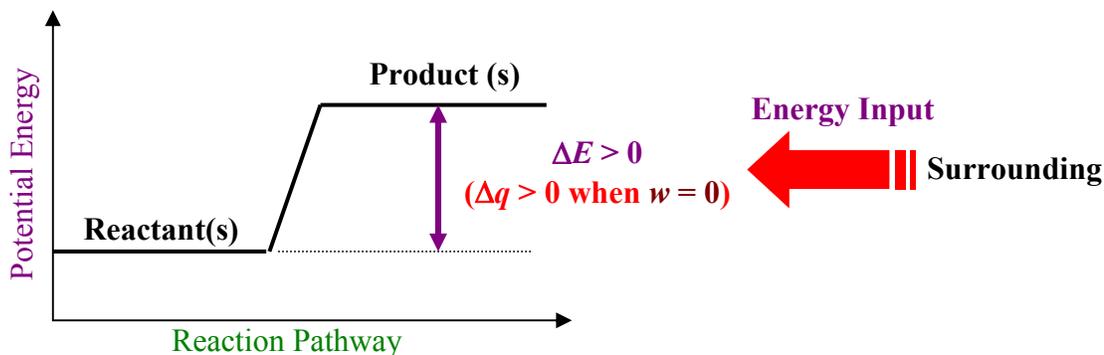
Exothermic Process ($\Delta E < 0$): - when energy flows “out” of the system into the surrounding.
(Surrounding gets Warmer.)

Potential Energy Diagram for Exothermic Process



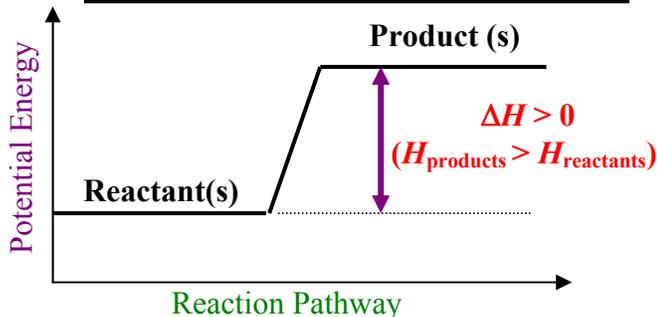
Endothermic Process ($\Delta E > 0$): - when energy flows into the system from the surrounding.
(Surrounding gets Colder.)

Potential Energy Diagram for Endothermic Process

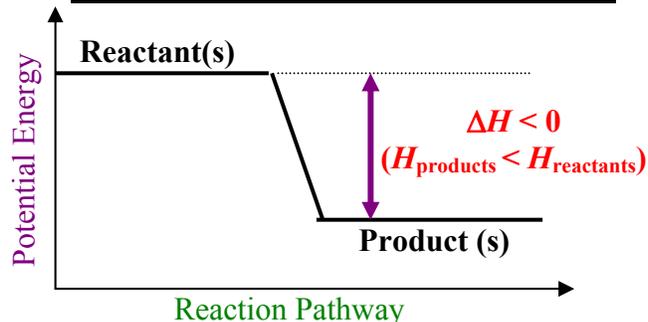


Molar Enthalpy Change (ΔH): - the amount of change in energy per mole of substance (J/mol or kJ/mol).

Potential Energy Diagram for Endothermic Chemical Reactions / Physical Processes



Potential Energy Diagram for Exothermic Chemical Reactions / Physical Processes



Writing ΔH Notations with Chemical Equations / Physical Process:

a. Endothermic Reactions / Processes



Example: Water is vaporized from its liquid state.



b. Exothermic Reactions / Processes



Example: Methane undergoes combustion at constant pressure.



Molar Enthalpy Change for Kinetic (Temperature) Change

$\Delta H = C\Delta T$ $\Delta T = T_f - T_i$

$\Delta H =$ Molar Enthalpy Change (J/mol kJ/mol) $\Delta H > 0$ (endothermic) $\Delta H < 0$ (exothermic)
 $\Delta T =$ Change in Temperature (in °C or K) $C =$ Molar Heat Capacity [J/(mol • °C or K)]

Example 1: Determine the molar enthalpy change in kJ/mol when water changes from 34.2°C to 79.6°C.

$C_{\text{water}} = 75.3 \text{ J/(mol} \cdot \text{°C)}$ $\Delta H = C\Delta T$
 $\Delta T = T_f - T_i = 79.6^\circ\text{C} - 34.2^\circ\text{C}$ $\Delta H = (75.3 \text{ J/(mol} \cdot \text{°C)})(45.4^\circ\text{C}) = 3418.62 \text{ J/mol}$
 $\Delta T = 45.4^\circ\text{C}$
 $\Delta H = ? \text{ kJ/mol}$ $\Delta H = 3.42 \text{ kJ/mol}$

Example 2: Calculate the change in temperature and the molar enthalpy change of iron when it cools from 243.7°C to 18.2°C. Comment on the nature of this thermodynamic process.

$C_{\text{iron}} = 25.1 \text{ J/(mol} \cdot \text{°C)}$ $\Delta H = C\Delta T$
 $\Delta T = T_f - T_i = 18.2^\circ\text{C} - 243.7^\circ\text{C}$ $\Delta H = (25.1 \text{ J/(mol} \cdot \text{°C)})(-225.5^\circ\text{C}) = -5660.05 \text{ J/mol}$
 $\Delta T = -225.5^\circ\text{C}$
 $\Delta H = ? \text{ kJ/mol}$ $\Delta H = -5.67 \text{ kJ/mol}$ (The process is exothermic;
 $\Delta H < 0$ because ΔT is negative - cooling)

Assignment

10.2 pg. 346 #1 to 3 (Practice); pg. 347 #1 to 3 (Practice); pg. 349 # 1 to 10

10.3: Changes in Enthalpy During Chemical Reactions

Standard State: - standard conditions of 1 atm and 25°C. It is denoted by a superscript “°”.

Standard Molar Enthalpy of Formation (ΔH_f°): - the amount of heat required / given off to make 1 mole of compound from its elemental components under standard conditions.

- the **Molar Heat of Formation of ALL ELEMENTS is 0 kJ.**
- the **state of the compound affects the magnitude of H_f .**
($H_2O_{(g)}$ has $\Delta H_f^\circ = -241.8$ kJ/mol; $H_2O_{(l)}$ has $\Delta H_f^\circ = -285.8$ kJ/mol)

(See Appendix A on pg. 833 to 834 in the Chemistry textbook for a list of ΔH_f°)

Standard Enthalpy of Formation (Chemical)

$$\Delta H = n\Delta H_f^\circ$$

ΔH = Change in Enthalpy n = moles ΔH_f° = Standard Molar Enthalpy of Formation (kJ/mol)

Example 1: Find the standard molar enthalpy of formation for table salt given that its formation reaction, $2 \text{Na}_{(s)} + \text{Cl}_{2(g)} \rightarrow 2 \text{NaCl}_{(s)} + 822 \text{ kJ}$, at standard conditions.

$$\begin{aligned} \Delta H &= -822 \text{ kJ} & \Delta H &= n\Delta H_f^\circ \\ n &= 2 \text{ mol of NaCl} & \Delta H_f^\circ &= \frac{\Delta H}{n} = \frac{-822 \text{ kJ}}{2 \text{ mol}} \\ \Delta H_f^\circ &= ? & & \Delta H_f^\circ = -411 \text{ kJ/mol} \end{aligned}$$

Example 2: What is the amount of heat absorbed / released when 100. g of $\text{CO}_{2(g)}$ is produced from its elements (CO_2 has $\Delta H_f^\circ = -393.5$ kJ/mol)?

$$\begin{aligned} n &= \frac{100. \text{ g}}{44.01 \text{ g/mol}} = 2.272210861 \text{ mol CO}_2 & \Delta H &= n\Delta H_f^\circ \\ & & \Delta H &= (2.272210861 \text{ mol})(-393.5 \text{ kJ/mol}) \\ \Delta H_f^\circ &= -393.5 \text{ kJ/mol} & & \Delta H = -894 \text{ kJ} \quad (894 \text{ kJ is released}) \\ \Delta H &= ? & & \end{aligned}$$

Example 3: Iron (III) oxide, rust, is produced from its elements, iron and oxygen. What is the mass of rust produced when 1.20 MJ is released when iron is reacted with oxygen ($\Delta H_f^\circ = -822.2$ kJ/mol for Fe_2O_3)?

$$\begin{aligned} \Delta H &= -1.20 \text{ MJ} = -1.20 \times 10^3 \text{ kJ (exothermic)} & \Delta H &= n\Delta H_f^\circ \\ \Delta H_f^\circ &= -822.2 \text{ kJ/mol} & n &= \frac{\Delta H}{\Delta H_f^\circ} = \frac{-1.20 \times 10^3 \text{ kJ}}{-822.2 \text{ kJ/mol}} = 1.459498905 \text{ mol} \\ M &= 159.7 \text{ g/mol Fe}_2\text{O}_3 & m &= nM = (1.459498905 \text{ mol})(159.7 \text{ g/mol}) \\ n &= ? & m &= 233 \text{ g} \end{aligned}$$

Example 4: Calculate the standard molar enthalpy of formation of silver (I) oxide when 91.2 g of Ag_2O is produced from its elements and 12.2 kJ of heat is released from the process.

$$\begin{aligned} n &= \frac{91.2 \text{ g}}{231.74 \text{ g/mol}} = 0.3935444895 \text{ mol Ag}_2\text{O} & \Delta H &= n\Delta H_f^\circ \\ \Delta H &= -12.2 \text{ kJ (exothermic)} & \Delta H_f^\circ &= \frac{\Delta H}{n} = \frac{-12.2 \text{ kJ}}{0.3935444895 \text{ mol}} \\ \Delta H_f^\circ &= ? & & \Delta H_f^\circ = -31.0 \text{ kJ/mol} \end{aligned}$$

Unit 6: Thermochemistry, States of Matter and Intermolecular Forces Chemistry

Molar Heat of Reaction (ΔH_{rxn}): - the amount of heat released when one mole of reactant undergoes various chemical or physical changes.
- examples are ΔH_{comb} , ΔH_{neut} (neutralization), ΔH_{sol} (solution).

Enthalpy of Chemical Reactions

$$\Delta H = n\Delta H_{rxn}$$

$$\Delta H = n\Delta H_{comb}$$

ΔH = Change in Enthalpy

n = moles

ΔH_{rxn} = Molar Heat of Reaction (kJ/mol)

ΔH_{comb} = Molar Heat of Combustion (kJ/mol)

Example 5: Given that $2 \text{C}_4\text{H}_{10(g)} + 13 \text{O}_{2(g)} \rightarrow 8 \text{CO}_{2(g)} + 10 \text{H}_2\text{O}_{(g)} + 5317 \text{ kJ}$, calculate the change in enthalpy when 28.2 g of butane is burned.

$$\Delta H_{rxn} = \frac{-5317 \text{ kJ}}{2 \text{ mol}} = -2658.5 \text{ kJ/mol}$$

(There are 2 moles of C_4H_{10} in the chemical equation for 5317 kJ.)

$$n = \frac{28.2 \text{ g}}{58.14 \text{ g/mol}} = 0.4850361197 \text{ mol C}_4\text{H}_{10}$$

$$\Delta H = n\Delta H_{rxn} = (0.4850361197 \text{ mol})(-2658.5 \text{ kJ/mol})$$

$\Delta H = ?$

$$\Delta H = -1.29 \times 10^3 \text{ kJ} = -1.29 \text{ MJ} \quad (1 \text{ MJ} = 1000 \text{ kJ})$$

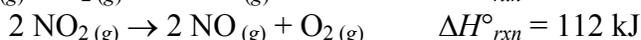
Standard Molar Enthalpy of Reaction (ΔH°_{rxn}): - the amount of heat involved when 1 mol of a particular product is produced or 1 mol of a particular reactant is consumed under standard conditions of 1 atm and 25°C.
(Note: In hydrocarbon combustion, assume all products are gaseous unless otherwise stated.)

Hess's Law: - the indirect method of obtaining overall ΔH°_{rxn} of a net reaction by the addition of ΔH°_{rxn} of a series of reactions.

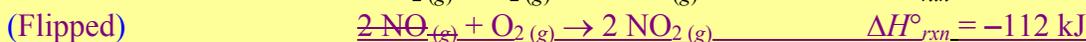
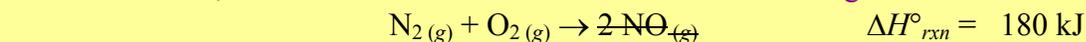
- when adding reactions, compare the reactants and products of the overall net reaction with the intermediate (step) reactions given. Decide on the intermediate reactions that need to be reversed and / or multiply by a coefficient, such that when added, the intermediate products will cancel out perfectly yielding the overall net reaction.
- if a particular reaction needs to be reversed (flipped), the sign of the ΔH for that reaction will also need to be reversed.
- if a coefficient is used to multiply a particular reaction, the ΔH for that reaction will also have to multiply by that same coefficient.

(Check out Hess's Law Animation at <http://intro.chem.okstate.edu/1314F00/Lecture/Chapter6/Hesslaw2.html>)

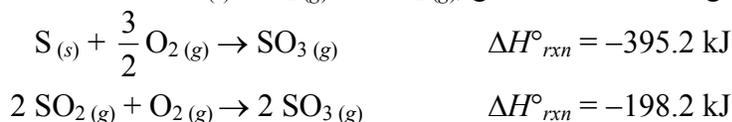
Example 6: Calculate ΔH°_{rxn} for the reaction $\text{N}_2(g) + 2 \text{O}_2(g) \rightarrow 2 \text{NO}_2(g)$, when given the following reactions.



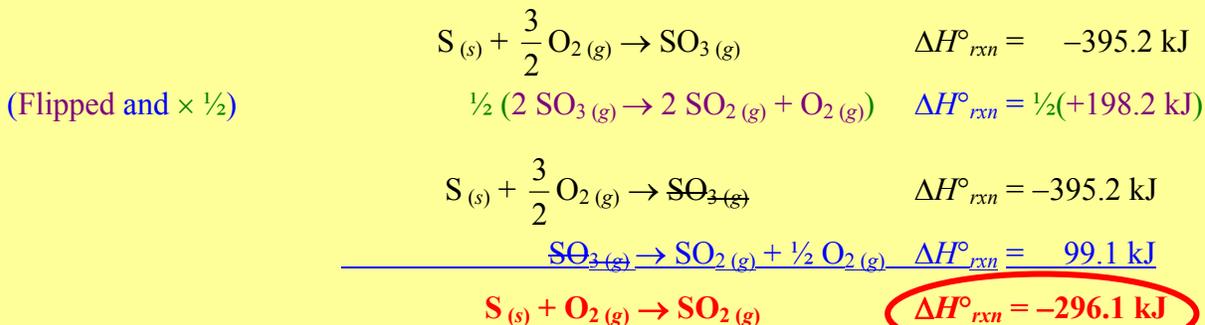
Note that 2 NO_2 in the net reaction is on the product side, whereas 2 NO_2 in the second reaction is on the reactant side. Hence, we need to reverse the second reaction and its sign of the ΔH°_{rxn} .



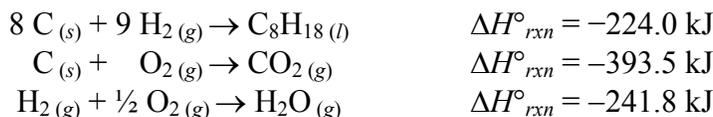
Example 7: Determine the ΔH°_{rxn} for the reaction $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$, given the following reactions.



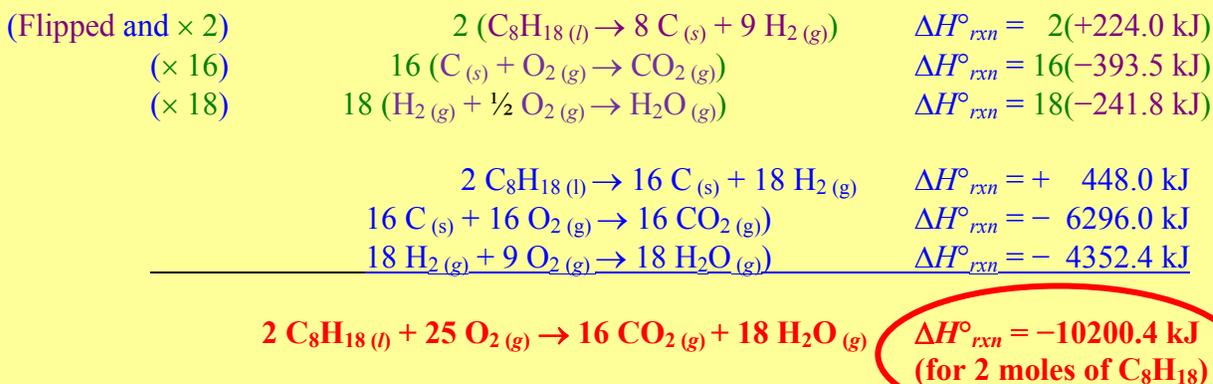
- SO_2 in the net reaction is on the product side, whereas 2 SO_2 in the second reaction is on the reactant side. Hence, we need to reverse the second reaction and its sign of the ΔH°_{rxn} .
- There is only 1 SO_2 in the net reaction, whereas there are 2 SO_2 in the second reaction. Therefore the second reaction and its ΔH°_{rxn} need to be multiply by the coefficient of $\frac{1}{2}$.



Example 8: Iso-octane is burned, $2 C_8H_{18(l)} + 25 O_{2(g)} \rightarrow 16 CO_{2(g)} + 18 H_2O_{(g)}$. Find the ΔH°_{rxn} for the overall reaction of when the following reactions are given.



- C_8H_{18} in the net reaction is on the reactant side; whereas there are 2 C_8H_{18} in the first reaction is on the product side. Hence, we need to reverse the first reaction (and its sign of the ΔH°_{rxn}), and multiply the first reaction (and its ΔH°_{rxn}) by the coefficient of 2.
- There are 16 CO_2 in the net reaction, whereas there is 1 CO_2 in the second reaction. Therefore the second reaction and its ΔH°_{rxn} need to be multiply by the coefficient of 16.
- There are 18 H_2O in the net reaction, whereas there is 2 H_2O in the third reaction. Therefore the third reaction and its ΔH°_{rxn} need to be multiply by the coefficient of 18.



$$\Delta H_{\text{comb}} = \frac{\Delta H}{n} = \frac{-10200.4 \text{ kJ}}{2 \text{ mol } C_8H_{18}}$$

$$\Delta H_{\text{comb}} = -5100.2 \text{ kJ/mol } C_8H_{18} \text{ burned}$$

Direct Method to determine Standard Enthalpy of Reaction:

- If the step-wise chemical equations are not given, there is no way we can use Hess's Law to find ΔH_{rxn} , then we must use a more direct method.
- it is equal to the difference between all enthalpies of products and all enthalpies of reactants.

Direct Method to determine Standard Enthalpy of Reaction

$$\Delta H^{\circ}_{rxn} = \Sigma H^{\circ}_{products} - \Sigma H^{\circ}_{reactants}$$

ΔH°_{rxn} = Change in Enthalpy of Reaction

$\Sigma H^{\circ}_{products}$ = Sum of Heat of Products (from all $n\Delta H^{\circ}_f$ of products)

$\Sigma H^{\circ}_{reactants}$ = Sum of Heat of Reactants (from all $n\Delta H^{\circ}_f$ of reactants)

Example 9: Find the amount of heat released when 34.9 g of butane gas is burned at standard conditions.
 ($\Delta H^{\circ}_f \text{C}_4\text{H}_{10} = -124.7 \text{ kJ/mol}$; $\Delta H^{\circ}_f \text{CO}_2 = -393.5 \text{ kJ/mol}$; $\Delta H^{\circ}_f \text{H}_2\text{O}_{(g)} = -241.8 \text{ kJ/mol}$)

We have to first write out a balance equation for the combustion of butane.



(We have to divide all coefficients by 2 because we are calculating ΔH_{rxn} per mol of butane burned.)



$$\Delta H_f: \quad -124.7 \text{ kJ/mol} \quad 0 \text{ kJ/mol} \quad -393.5 \text{ kJ/mol} \quad -241.8 \text{ kJ/mol}$$

$$\Delta H^{\circ}_{rxn} = \Sigma H^{\circ}_{products} - \Sigma H^{\circ}_{reactants}$$

$$\Delta H^{\circ}_{rxn} = [4 \text{ mol} (-393.5 \text{ kJ/mol}) + 5 \text{ mol} (-241.8 \text{ kJ/mol})] - [1 \text{ mol} (-124.7 \text{ kJ/mol}) + \frac{13}{2} \text{ mol} (0 \text{ kJ/mol})]$$

$$\Delta H^{\circ}_{rxn} = [-2783 \text{ kJ}] - [-124.7 \text{ kJ}]$$

$$\Delta H_{rxn} = -2658.3 \text{ kJ/mol of C}_4\text{H}_{10} \text{ burned}$$

$$n = \frac{34.9 \text{ g}}{58.14 \text{ g/mol}} = 0.6002751978 \text{ mol C}_4\text{H}_{10}$$

$$\Delta H = n\Delta H_{rxn}$$

$$\Delta H = (0.6002751978 \text{ mol})(-2658.3 \text{ kJ/mol}) = -1595.7 \text{ kJ}$$

$$\Delta H_{rxn} = -2658.3 \text{ kJ/mol}$$

$$\Delta H = ?$$

$$\Delta H = -1.60 \text{ MJ} \quad (1.60 \text{ MJ is released})$$

Assignment

10.3 (Part 1) pg. 356 #1 & 2 (Practice); pg. 357 #1 & 2 (Practice)

Constant-Pressure Calorimeter (or Styrofoam Calorimeter) - commonly used to determine ΔH_{neut} , ΔH_{ion} ,

A Simple Styrofoam Calorimeter



ΔH_{fus} , ΔH_{vap} , ΔH_{rxn} of non-combustion reaction. First, the sample's mass is measured. Water is commonly used to absorb or provide the heat for the necessary change. The initial and final temperatures of the water are recorded, allowing us to find the amount of heat change. By applying the law of conservation of energy, we can then calculate the necessary molar enthalpy of change.

Chemistry Unit 6: Thermochemistry, States of Matter and Intermolecular Forces

First Law of Thermodynamics: - states that energy cannot be created or destroyed. It can only be converted from one form to another. Therefore, energy in the universe is a constant
- also known as the **Law of Conservation of Energy** ($\Sigma E_{\text{initial}} = \Sigma E_{\text{final}}$).

Calorimetry: - uses the conservation of energy (**Heat Gained = Heat Lost**) to measure calories (old unit of heat: 1 cal = 4.184 J).
- physical calorimetry involves the mixing of two systems (one hotter than the other) to reach some final temperature.
- **the key to do these problems is to identify which system is gaining heat and which one is losing heat.**

Example 10: Hot water at 90.0°C is poured into 100. g of cold water at 10.0°C. The final temperature of the mixture is 70.0°C. Determine the mass of the hot water.

$$m_{\text{cold water}} = 100. \text{ g}$$

$$c_p = 4.18 \text{ J/(g} \cdot \text{ }^\circ\text{C)}$$

$$\Delta T_{\text{hot water}} = 90.0^\circ\text{C} - 70.0^\circ\text{C} = 20.0^\circ\text{C}$$

$$\Delta T_{\text{cold water}} = 70.0^\circ\text{C} - 10.0^\circ\text{C} = 60.0^\circ\text{C}$$

$$m_{\text{hot water}} = ?$$

Heat Lost = Heat Gained
(hot water, kinetic) (cold water, kinetic)

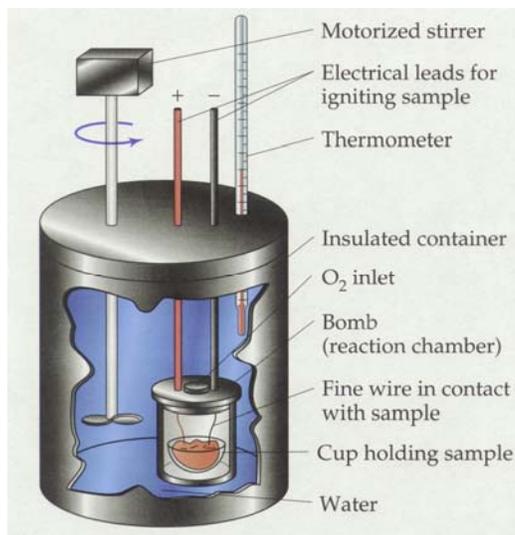
$$m_{\text{hw}} c_p \Delta T_{\text{hw}} = m_{\text{cw}} c_p \Delta T_{\text{cw}}$$

$$m_{\text{hw}} = \frac{m_{\text{cw}} \Delta T_{\text{cw}}}{\Delta T_{\text{hw}}} = \frac{(100. \text{ g})(60.0^\circ\text{C})}{(20.0^\circ\text{C})}$$

$$m_{\text{hw}} = 300. \text{ g}$$

Molar Heat of Combustion (ΔH_{comb}): - the amount of heat released when one mole of reactant is burned with excess oxygen.

- the reaction is often exothermic and therefore $\Delta H_{\text{comb}} < 0$.



Schematic of a Bomb Calorimeter

Enthalpy of Combustion

$$\Delta H = n \Delta H_{\text{comb}}$$

ΔH = Change in Enthalpy n = moles

ΔH_{comb} = Molar Heat of Combustion (kJ/mol)

- we often use a *constant-volume calorimeter* (or *bomb calorimeter*) to determine ΔH_{comb} due to its well-insulated design. It is calibrated for the heat capacity of the calorimeter, C_{cal} , before being used to calculate ΔH_{comb} of other substances. The sample is measured and burned using an electrical ignition device. Water is commonly used to absorb the heat generated by the reaction. The temperature of the water increases, allowing us to find the amount of heat generated. By applying the law of conservation of energy, we can then calculate the ΔH_{comb} of the sample.

Chemical Combustion Calorimetry

Heat Lost = Heat Gained
(Combustion Reaction) (water, kinetic)

$$n_{\text{sample}} \Delta H_{\text{comb}} = C_{\text{cal}} \Delta T \text{ (if bomb calorimeter is used)}$$

or

$$n_{\text{sample}} \Delta H_{\text{comb}} = m_w c_{p,w} \Delta T \text{ (if the heat absorbed by the calorimeter itself is ignored)}$$

Example 11: Octane, $C_8H_{18(l)}$ was burned completely to $CO_2(g)$ and $H_2O(l)$ in a bomb calorimeter. The following are the observations of the experiment.

Mass of $C_8H_{18(l)}$ burned	32.65 g
Initial Temperature of Calorimeter and Water	16.50°C
Final Temperature of Calorimeter and Water	77.30°C
Heat Capacity of Calorimeter	22.70 kJ/°C

- Determine the experimental molar heat of combustion of $C_8H_{18(l)}$.
- The theoretical ΔH_{comb} for iso-octane ($C_8H_{18(l)}$) is -5100.2 kJ/mol (from Example 8), calculate the % error of this experiment.

a. $m_{octane} = 32.65$ g
 $n_{octane} = \frac{32.65 \text{ g}}{114.26 \text{ g/mol}}$
 $= 0.2857517942$ mol C_8H_{18}

$C_{cal} = 22.70$ kJ/°C
 $\Delta T = T_f - T_i = 77.30^\circ\text{C} - 16.50^\circ\text{C}$
 $\Delta T = 60.80^\circ\text{C}$

$\Delta H_{comb} = ?$

Heat Lost = Heat Gained
 (Combustion Reaction) (water, kinetic)

$n_{octane}\Delta H_{comb} = C_{cal}\Delta T$

$\Delta H_{comb} = \frac{C_{cal}\Delta T}{n_{octane}} = \frac{(22.70 \text{ kJ/}^\circ\text{C})(60.80^\circ\text{C})}{(0.2857517942 \text{ mol})}$

$\Delta H_{comb} = -4830$ kJ/mol

(negative sign is added because ΔH_{comb} is exothermic)

b. % error = $\left| \frac{\text{Theoretical} - \text{Experimental}}{\text{Theoretical}} \right| \times 100\%$

% error = $\left| \frac{(-5100.2 \text{ kJ/mol}) - (-4830 \text{ kJ/mol})}{(-5100.2 \text{ kJ/mol})} \right| \times 100\%$

% error = 5.298%

(The small % error means the bomb calorimeter was a good heat insulator.)

Example 12: When 10.02 g of liquid heptane (C_7H_{16}) is burned in the reaction vessel of a calorimeter, 1.50 L of water around the vessel increased its temperature from 20.0°C to 85.0°C. Ignoring the metallic material of the calorimeter,

- determine the experimental standard molar enthalpy of combustion heptane.
- find the theoretical standard molar enthalpy of combustion of heptane.
 (ΔH_f° heptane = -224.2 kJ/mol C_7H_{16})
- draw its potential energy diagram.
- explain why the experimental ΔH_{rxn}° is different than its theoretical counterpart.

a. We use the conservation of heat to calculate experimental ΔH_{rxn} .

$n = \frac{10.02 \text{ g}}{100.23 \text{ g/mol}}$
 $n = 0.0999700688$ mol C_7H_{16}

$m_{water} = 1.50$ kg (1 kg = 1 L of water)
 $\Delta T = 85.0^\circ\text{C} - 20.0^\circ\text{C} = 65.0^\circ\text{C}$
 $c_p = 4.18$ kJ / (kg • °C)

$\Delta H_{rxn}^\circ = ?$

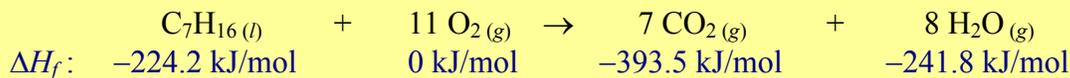
$n\Delta H_{rxn}^\circ = m_w c_p \Delta T$

$\Delta H_{rxn}^\circ = \frac{m_w c_p \Delta T}{n} = \frac{(1.50 \text{ kg})(4.18 \text{ kJ/(kg} \cdot ^\circ\text{C)})(65.0^\circ\text{C})}{(0.0999700688 \text{ mol})}$

$\Delta H_{rxn}^\circ = 4076.72021$ kJ/mol (released)

Experimental $\Delta H_{rxn}^\circ = -4.08$ MJ/mol of C_7H_{16} burned

b. To find theoretical ΔH°_{rxn} for the combustion of heptane, we have to use the direct method.



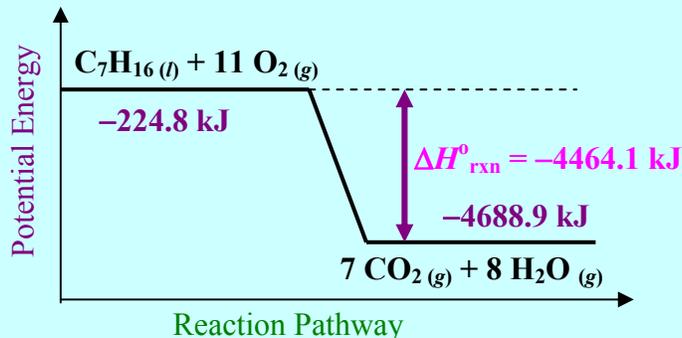
$$\Delta H^\circ_{rxn} = \Sigma H^\circ_{\text{products}} - \Sigma H^\circ_{\text{reactants}}$$

$$\Delta H^\circ_{rxn} = [7 \text{ mol} (-393.5 \text{ kJ/mol}) + 8 \text{ mol} (-241.8 \text{ kJ/mol})] - [1 \text{ mol} (-224.8 \text{ kJ/mol})]$$

$$\Delta H^\circ_{rxn} = [-4688.9 \text{ kJ}] - [-224.8 \text{ kJ}] = -4464.1 \text{ kJ}$$

Theoretical $\Delta H^\circ_{rxn} = -4.46 \text{ MJ/mol}$ of C_7H_{16} burned

c. Potential Energy Diagram



d. Some of the possible reasons why experimental ΔH°_{rxn} (-4.08 MJ) is different than the theoretical ΔH°_{rxn} (-4.46 MJ)

- **Some of the heat released by the reaction is absorbed by the metal calorimeter itself. Thus, the temperature gained by the water is not an exact reflection of the energy lost by the combustion.**
- **The calorimeter is not a closed system. Heat might escape into the surrounding.**
- **Even if the system is closed, the buildup of gases from the reaction would increase pressure and volume. Hence, some of the energy produced from the reaction is used to do work by the system. Thereby, lowering the heat available to warm the water.**

Assignment

10.3 (Part 2) pg. 357 #1 to 7

Chapter 10 Review pg. 370–372 #22 to 29; 36, 38, 40, 43

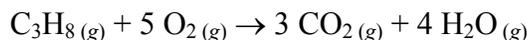
(Extra Section): Present Sources of Energy and New Energy Sources

Fossil Fuel: - hydrocarbon fuels that came from fossils of decayed organisms.

1. **Natural Gas:** - fossil fuel that consists of mainly small alkanes (80% methane, 10% ethane, 4% propane, 2% butane, 4% nitrogen).
- usually burns efficiently (complete combustion).

Complete Combustion: - where the products of combustion are carbon dioxide and water vapour only.
-characterized by a blue flame.

Example: Propane burns completely.



Incomplete Combustion: - where the main product of combustion is carbon monoxide, along with carbon dioxide and water vapour.
- happens when carbon particles started to form during combustion and deposited as soot as they cooled, or when there is insufficient oxygen.
- characterized by a yellow flame.

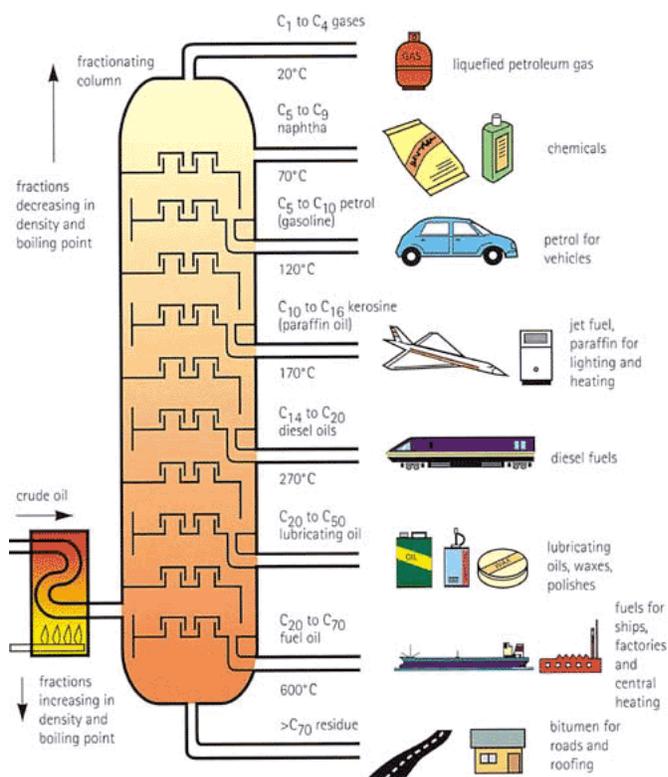
Example: Incomplete combustion of Propane.



2. **Petroleum (Crude Oil):** - fossil fuels that consist mainly of heavier alkanes along with small amounts of aromatic hydrocarbons, and organic compounds that contain sulfur, oxygen and nitrogen.
- gasoline is composed of 40% of crude oil, whereas natural gas is composed of only 10%.

Fractional Distillation: - a method of heating crude oil in a tall column to separate its different components by their different boiling points.

- lighter alkanes in the natural gas will rise up to the top of the column because of their low boiling points.
- the heavier, fuel and lubricating oils will boil off at the bottom of the column due to their high boiling points.



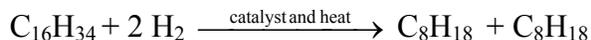
Petroleum Refining: - a process to isolate different types of fuel from crude oil using fractional distillation or cracking.

Chemistry Unit 6: Thermochemistry, States of Matter and Intermolecular Forces

Cracking: - a chemical process whereby bigger alkanes are broken up into smaller ones using a catalyst and heat.

- since gasoline and natural gas only consists of 50% of crude oil, cracking is necessary to convert heavier fuel to more common fuel used in today's world.

Example: The Cracking of Hexadecane.



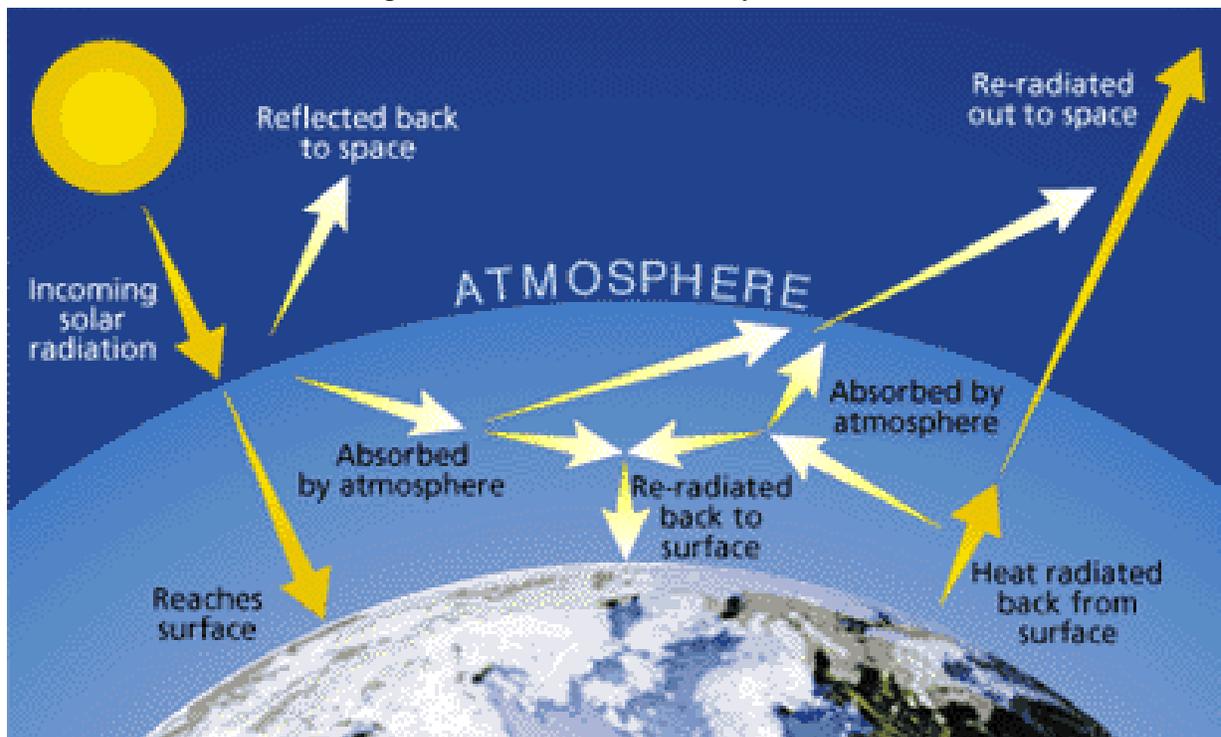
Reforming: - a chemical process where smaller alkanes are combined together and hydrogen is removed to form heavier alkanes or changed unbranched alkanes into branched alkanes.

- branched alkanes are easier to burn and has a higher octane value in gasoline. (isooctane or 2,2,4-trimethylpentane has the best octane rating – assigned as 100)

3. **Coal:** - a carbon-based mineral consists of very dense hydrocarbon ring compounds with high molar masses.

- leaves a lot of soot and burns incompletely.
- usually contains 7% sulfur and when combusted with oxygen gives off SO_2 and SO_3 , which is the main source of air pollution and acid rain.

Greenhouse Effect: - the emission of greenhouses gases that traps more of the sun's radiant (heat) energy in the atmosphere than it occurs naturally.



Greenhouses Gases: - man-made and naturally occur gases that contribute to the Greenhouse Effect.

1. **Carbon dioxide (CO_2)** is released to the atmosphere when solid waste, fossil fuels (oil, natural gas, and coal), and wood and wood products are burned.

- Methane (CH₄)** is emitted during the production and transport of coal, natural gas, and oil. Methane emissions also result from the decomposition of organic wastes in municipal solid waste landfills, and the raising of livestock.
- Nitrous oxide (N₂O)** is emitted during agricultural and industrial activities, as well as during combustion of solid waste and fossil fuels.
- Hydrofluorocarbons (HFCs), Perfluorocarbons (PFCs), and Sulfur Hexafluoride (SF₆)** are very powerful greenhouse gases that are not naturally occurring that are generated in a variety of industrial processes.

Each greenhouse gas differs in its ability to absorb heat in the atmosphere. HFCs and PFCs are the most heat-absorbent. Methane traps over 21 times more heat per molecule than carbon dioxide, and nitrous oxide absorbs 270 times more heat per molecule than carbon dioxide. Often, estimates of greenhouse gas emissions are presented in units of millions of metric tons of carbon equivalents (MMTCE), which weights each gas by its GWP value, or Global Warming Potential. (Information from US. EPA)

- Automobiles and Major Transportations account for 34% of CO₂ emissions globally (Power Plants contributes 33%; Major Industries and Home Heating contribute the remaining 33%).
- Presently 89% of Energy Productions involve the burning of Fossil Fuels (Coal, Petroleum, Natural Gas and Biomass).
- Heat and Electricity generated from combustion of fossil fuel is at most 30% efficient.

(Data from University of Michigan: <http://www.umich.edu/~gs265/society/greenhouse.htm>)

The Environmental Effect of Using Fossil Fuel: (Greenhouse Effect)

- Global Warming**: - the warming of global temperature due to an increased of greenhouse gases in the atmosphere.
- Rise of Water Level**: - low-lying islands and coastal area are endangered as polar icecaps melt due to the rise of temperature as a result of the greenhouse effect.
- Unpredicted and Erratic Climate**: - greenhouse effect is related to droughts and dry whether in many parts of the world.
- Deforestation**: - another cause of an increased in CO₂ level in the atmosphere. As forests disappeared, there is a lack of plants to absorb carbon dioxide using photosynthesis.
- also causes mud and landslides, demineralization of the soil, lost animal habitats and extinction, destruction of entire ecosystems. Plants that may have important medicinal values can also be destroyed.

Alternate Energy Sources without the Emission of Greenhouse Gas

- Solar Energy**: - the most efficient energy source where energy from the sun is converted directly to electricity through the use of photovoltaic cells (solar panels) or heat using high efficient insulated glass and an effective water heating system.
- technology exists but fairly expensive; requires many solar panels to generate adequate amount of electricity.

2. **Wind Power**: - the use of wind turbines to generate electricity.
 - very efficient and extremely economical, but location specific and not very reliable when there is no wind.
 - can disrupt migratory routes of birds (they get caught in the turbine), aesthetic problems for the various landscapes.

3. **Geothermal Power**: - the use of underground steam to generate electricity.
 - very efficient and reliable, but location specific.
 - geothermal power is widely use in Iceland where it is sitting on the Atlantic ridge and there are lots of hot springs.

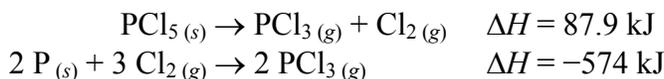
4. **Tidal Power**: - the use of tidal current to generate electricity.
 - very efficient and somewhat reliable, but location specific.
 - tidal power involves placing electric turbines at a narrow mouth of a channel where normal tides can cause bigger current and quick rise in water levels. It is being used in the Bay of Fundy at Nova Scotia, Canada and Kvalsund at the Arctic tip of Norway.
 - tidal power can sometimes disrupt migratory routes of marine species.

5. **Hydroelectricity**: - the use of dam and reservoir to turn electric turbines as water falls from a higher level to the spillway (potential energy converted to kinetic energy to electricity).
 - very efficient and no emission of greenhouse gas.
 - location specific and very expensive to built. The reservoir flooding can destroy ecological habitats and force migrations of people from towns and villages (Aswan Dam in Egypt and the Three Gorges Dam in China displaced thousands of people and submerged ancient cities). The presence of the dam can disrupt aquatic migratory routes as well.
 - dams have a limited life span (the collection of silt and mud at the bottom of the dam has to be clear periodically to maintain the structural integrity of the dam). Dams can burst during earthquakes or poor maintenance. Flash flooding of towns along spillway is always a danger.

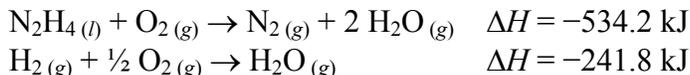
6. **Hydrogen Fuel**: - burning hydrogen to form water and generate heat and electricity.
 - very efficient and zero pollution.
 - hydrogen is very explosive and technologies are still needed for supplying and storing hydrogen safely in automobiles and homes.

Worksheet: ΔH , Hess's Law and Calorimetry

- How much heat does 32.0 g of water absorb when it is heated from 25.0°C to 90.0°C?
- When 525 J of heat is added to 3.40 g of olive oil at 20.0°C, the temperature increases to 85.0°C. What is the specific heat of olive oil?
- A piece of silver has a molar heat capacity of 25.35 J/(mol • °C). What is the enthalpy change when 3.00 mol of silver is heated from 30.0°C to 140°C?
- Ethanol (C₂H₅OH (l)) has a molar enthalpy of 112.4 J/(mol • °C). If 6.86 kJ of heat is removed from 5.65 mol of ethanol from an initial temperature of 68.0°C, what is its final temperature?
- What is the amount of heat released / absorbed when 23.5 g of LiCl (s) is formed from its elements?
- Determine the amount of heat released / absorbed when 50.2 g of SiCl₄ (s) is decomposed into its elements.
- The combustion of ethane, C₂H₄ (g) + 3 O₂ (g) → 2 CO₂ (g) + 2 H₂O (l) + 1390 kJ, is an exothermic reaction. How much heat is given off when 75.9 g of C₂H₄ is burned?
- The combustion of decane, 2 C₁₀H₂₂ (l) + 31 O₂ (g) → 20 CO₂ (g) + 22 H₂O (l) + 13.6 MJ, is an exothermic reaction. How much heat is given off when 75.9 g of C₁₀H₂₂ is burned?
- The formation reaction of phosphorus pentachloride has an equation, 2 P (s) + 5 Cl₂ (g) → 2 PCl₅ (s). Find the enthalpy change by using the following thermochemical equations.



- Given the following reactions,



Calculate the heat of formation of N₂H₄ (l), which has an equation N₂ (g) + 2 H₂ (g) → N₂H₄ (l).

- A 150.0 g sample of a metal at 75.0°C is added to 150.0 g of water at 15.0°C. The temperature of the water and metal reach a final temperature of 18.3°C. Calculate the specific heat of the metal in J/(g • °C).
- A 0.1964 g sample of quinine (C₆H₄O₂) is burned in a bomb calorimeter that has a heat capacity of 1.56 kJ/°C. The temperature of the calorimeter increases by 3.20°C. Calculate the molar enthalpy of combustion of quinine in kJ/mol.
- The following calorimetry experiment yields the following results.

Mass of water used	700 g
Mass of pentane (C ₅ H ₁₂) used	4.34 g
Initial Temperature of Water	15.8°C
Final Temperature of Water	82.1°C

Determine the experimental molar heat of combustion of pentane (C₅H₁₂).

Answers

- | | | |
|---|------------------------------------|------------------------------------|
| 1. 8.69 kJ | 2. 2.38 J/(g • °C) | 3. 8.37 kJ |
| 4. 57.2°C | 5. 227 kJ released | 6. 194 kJ absorbed |
| 7. 3.76 × 10 ³ kJ or 3.76 MJ | 8. 3.63 MJ | 9. -749.8 kJ |
| 11. 0.243 J/(g • °C) | 12. -2.75 × 10 ³ kJ/mol | 13. -3.23 × 10 ³ kJ/mol |

Chapter 11: States of Matter and Intermolecular Forces

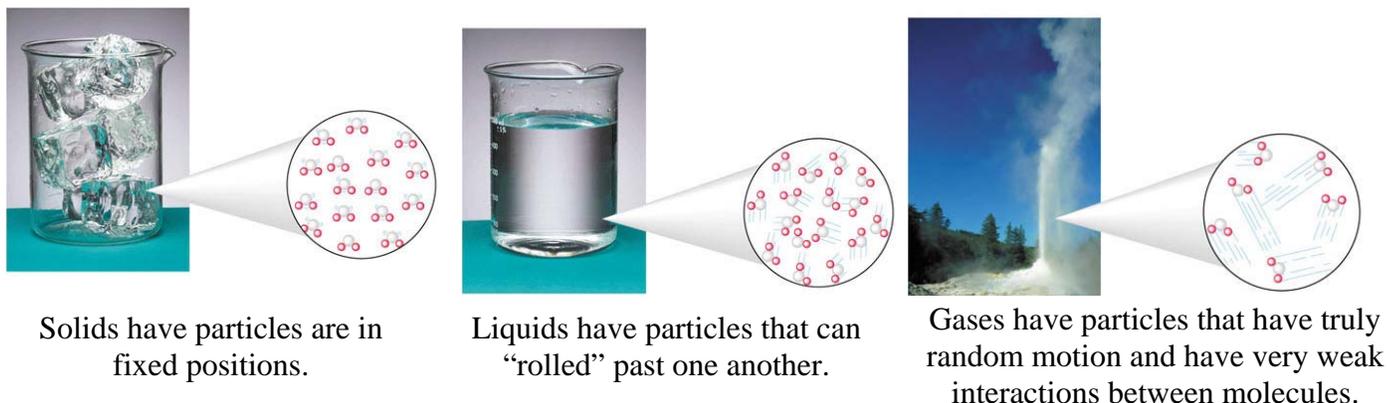
11.1 & 11.2: States and State Changes & Intermolecular Forces

Solids: - the state of matter where it has a definite volume with a constant shape.
- particles in solid states have fixed positions and can only vibrate in one place.

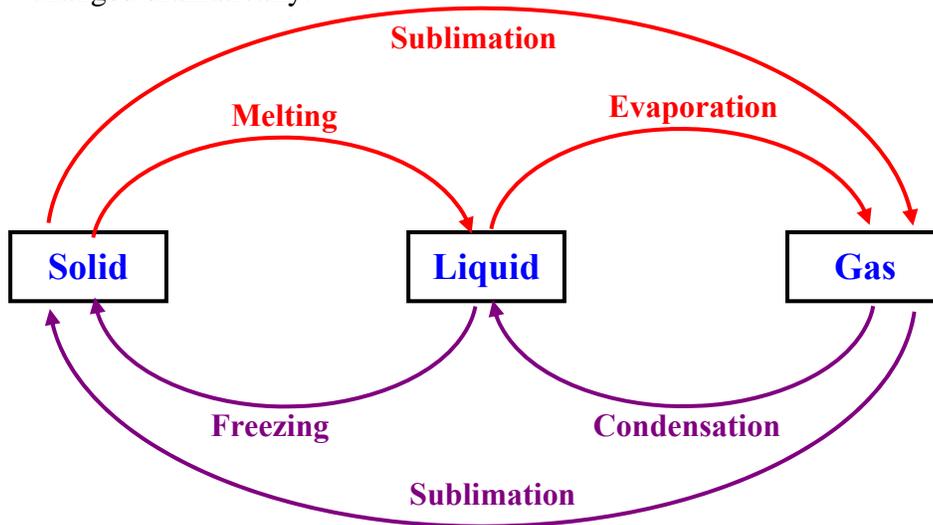
Liquid: - the state of matter where it has a definite volume but an indefinite shape.
- particles do not have fixed positions, and they can move past each other. The ease in which they can slide past each other depends on the forces between particles (see next section).

Gas: - the state of matter where it has an indefinite volume and shape (compressible).
- particles in gas phase do not have any attractive forces between each other. They are constantly moving and they have a large amount of space between themselves.

Kinetic Molecular Theory and States of Matter



Phase Change: - when particles gained enough energy, such that their motions relative to each other are changed dramatically.



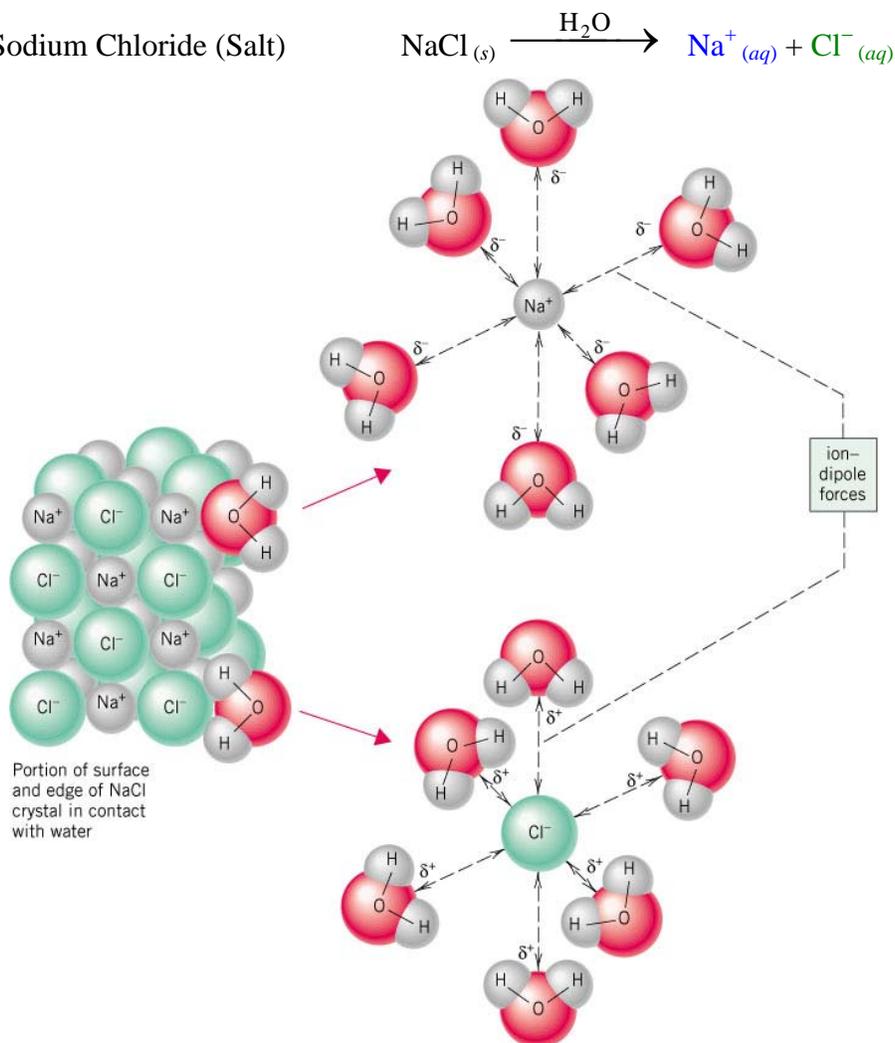
Boiling Point: - the temperature at which a substance's liquid phase change into its gaseous phase at a given pressure. Different pressures might affect boiling points. At this temperature, the particles gained enough energy to overcome the attractive forces between each other. Both the liquid and gaseous phases can coexist at this temperature.

Melting Point: - the temperature at which a substance's solid phase change into its liquid phase at a given pressure. Different pressures might affect boiling points. At this temperature, the particles gained enough energy to loosen from their fixed positions and can glide past each other. Both the solid and liquid phases can coexist at this temperature.
 - also known as the **freezing point**.

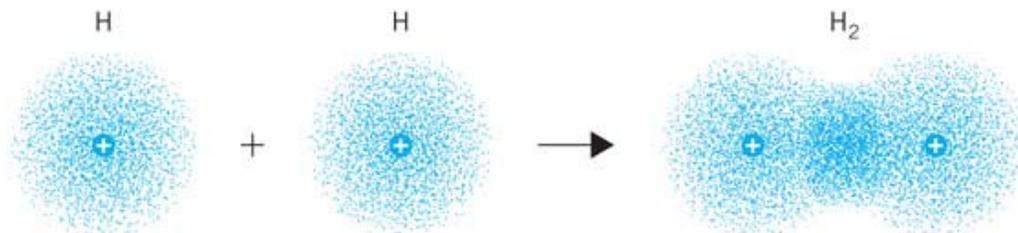
Intermolecular Forces: - attraction forces between molecules in a compound
 - the strengths of the intermolecular forces **explain** the **physical properties of compounds (solubility, boiling and freezing points)**.

- a. **Ion-Dipole Attraction:** - when ionic compounds dissolve in water, the cation and anion components separate from one another. These ions are then attracted by the polar water molecules.
 - ion-dipole attraction is the **STRONGEST of all intermolecular forces**.

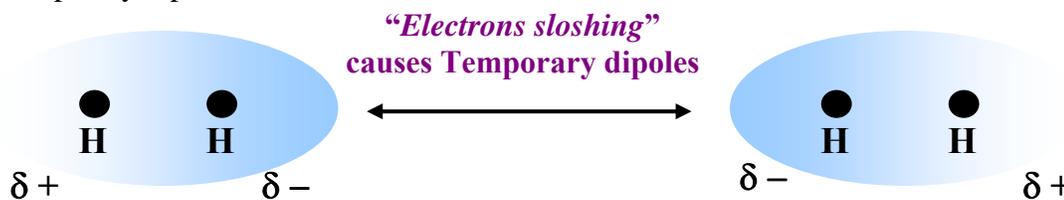
Example 1: Sodium Chloride (Salt)



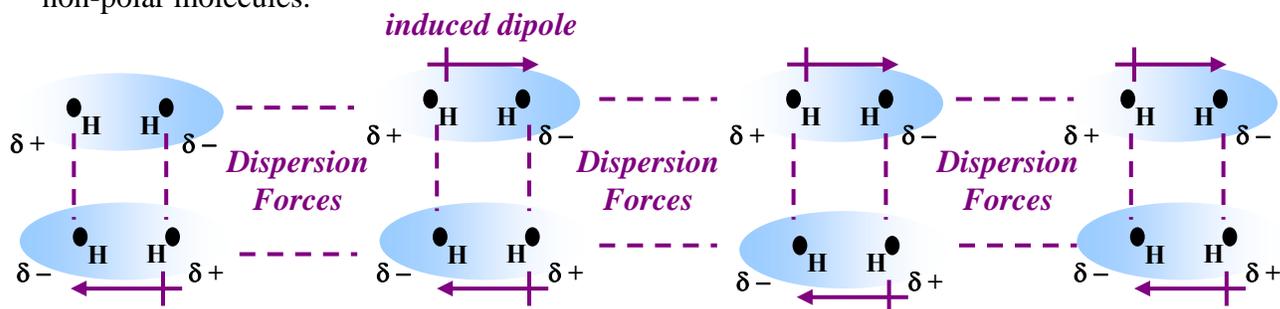
- b. **van der Waals Forces**: - named after Johannes van der Waals who studied real gases and interaction between molecules.
- there are two kinds of van der Waals forces.
 - they are **Dispersion Forces** and **Dipole-Dipole Interactions**.
- i. **Dispersion Forces**: - also known as **London Dispersion Forces** (named after Fritz London who first proposed how this force works).
- on average, the non-polar molecules do not have any permanent dipoles like polar molecules.



- the “*dispersion*” is the **temporary dipole** that forms within the molecules even in non-polar molecules due to the constant motions of electrons. In one instance, they can move to one side of the molecule making it temporary polar. In another instance, electrons will move and the direction of this temporary dipole will switch.



- This constant “sloshing around” of electrons causes non-polar molecules to have these **temporary dipoles**. These temporary “*induced*” dipoles are what cause the attractions between non-polar molecules.



- even monoatomic element like Helium has London Forces. (Check out animation at <http://virtuallaboratory.colorado.edu/Chemistry/StudentMaterials/applet%20pages/applets/london%20animation.gif>)
- in general, **the more electrons there are in a molecule, the stronger the London Dispersion Force** (attraction between molecules – intermolecular force). **This causes an increase in melting and boiling points of the chemical.**
- Note: All molecules have electrons. Hence, ALL molecules have London Dispersion Force.**

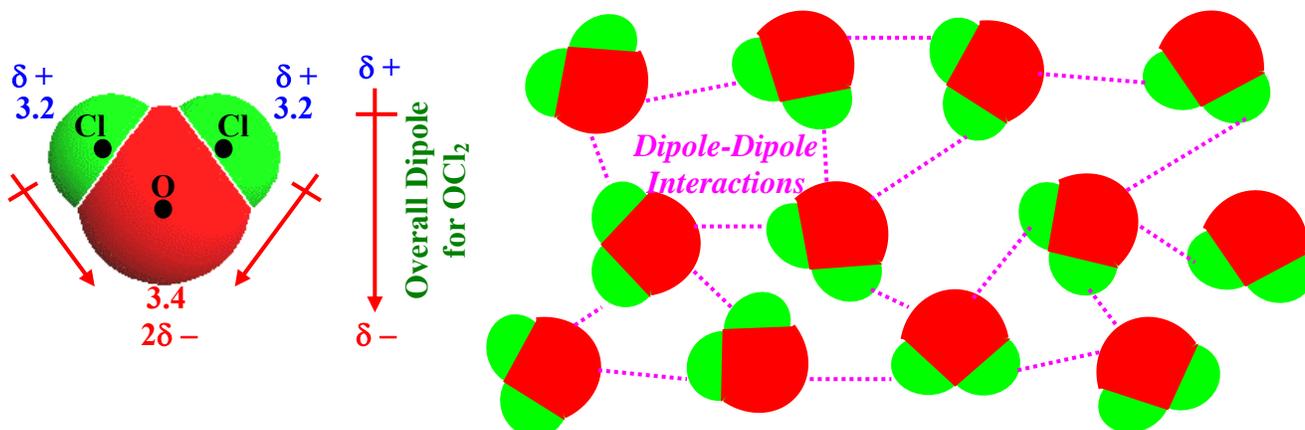
of electrons in atom or molecule ↑, London Dispersion Force ↑, Melting and Boiling Point ↑

Example 1: Explain the boiling points and the melting points of the noble gases.

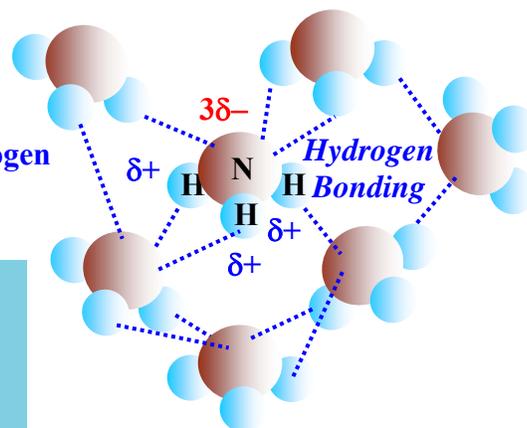
Noble Gases	# of e^-	Melting Point	Boiling Point
He	2	-272°C (1 K)	-269°C (4 K)
Ne	10	-249°C (24 K)	-246°C (27 K)
Ar	18	-189°C (84 K)	-186°C (87 K)
Kr	36	-157°C (116 K)	-153°C (120 K)
Xe	54	-112°C (161 K)	-108°C (165 K)
Rn	86	-71°C (202 K)	-62°C (211 K)

All atoms of noble gases are monoatomic non-polar. The only intermolecular force that governs the melting and boiling points is the London Dispersion Force. As the number of electrons in the noble gases increase, London dispersion force makes the attraction between the atoms greater. This in turn has an effect of increasing the boiling and melting point of the noble gas as one goes down the column.

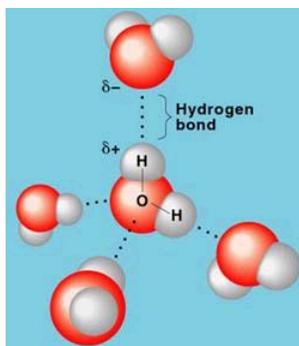
- ii. **Dipole-Dipole Interactions:** - also known as simply **Dipole Interactions**.
 - intermolecular forces resulted from polar molecules.
 - dipole interaction is **much stronger than Dispersion Force**.



- c. **Hydrogen Bonds:** - are intermolecular bonds that involve hydrogen atom with very electronegative atom that also consists of lone pairs.
 - these include **O-H, N-H, H-Cl and H-F bonds**.
 - the resulting molecule is always polar. Therefore, **all hydrogen bonding molecules also have dipole interactions**.
 - hydrogen bond is the **STRONGEST** of the covalent intermolecular forces.



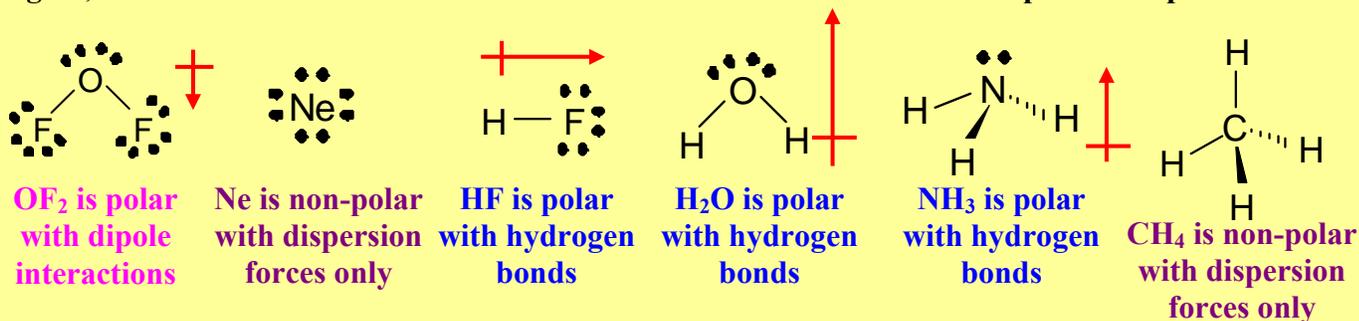
(View the Hydrogen Bond Animation at <http://www.northland.cc.mn.us/biology/Biology111/animations/hydrogenbonds>)



Example 2: Account for the differences in the boiling points of the compounds listed below.

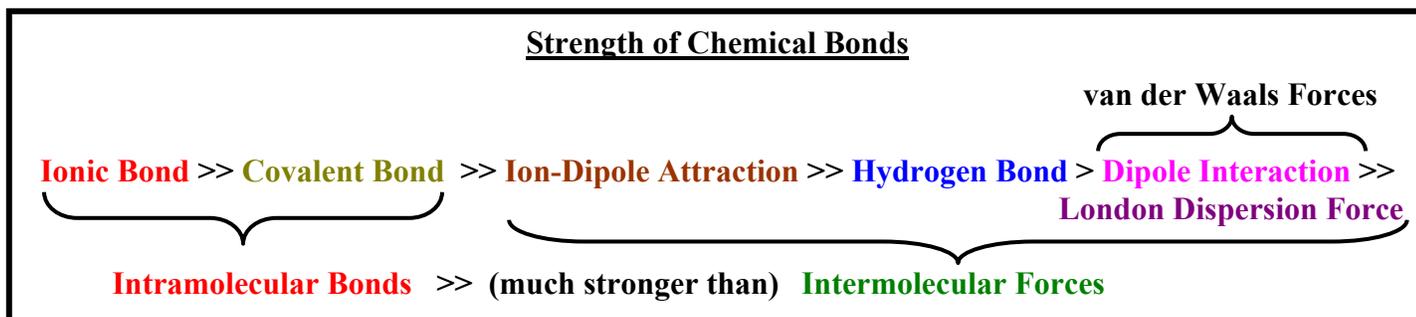
Molecule	Number of Electrons	London Dispersion Forces	Dipole Interactions	Hydrogen Bonds	Boiling Point
OF ₂	26	✓	✓	✗	-145°C (128 K)
Ne	10	✓	✗	✗	-246°C (27 K)
HF	10	✓	✓	✓	19°C (292 K)
H ₂ O	10	✓	✓	✓	100°C (373 K)
NH ₃	10	✓	✓	✓	-33°C (240 K)
CH ₄	10	✓	✗	✗	-161°C (112 K)

Again, we need to draw the structural formulas of these molecules and compare their polarities.



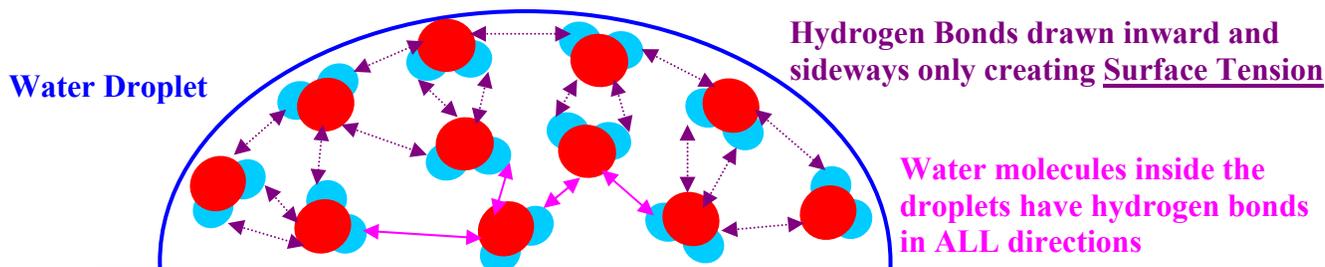
Summary of Chemical Bonding

- Intramolecular Bonds** such as covalent and ionic bonds deal with bond strength between atoms within an atom. They explain amount of energy in a **chemical change**.
- Intermolecular Forces** involve in a compound explain its **physical properties** such as solubility (“like dissolves like” – section 7.3, boiling and melting points – energy involved in physical phase change).
- van der Waals Forces** consist of **London Dispersion forces** (apply to all molecules) and **Dipole Interactions** (apply to polar molecules).
- Hydrogen Bonding** is the *strongest* of all **covalent intermolecular forces**.
- Ion-Dipole Attraction** is the *strongest* of all **intermolecular forces**.



Surface Tension: - the inward force of a liquid to minimize its surface area.

- water's hydrogen bond cannot attract with molecules in the air. Therefore, the higher net force is pushed inward, leaving a spherical surface with a high surface tension.

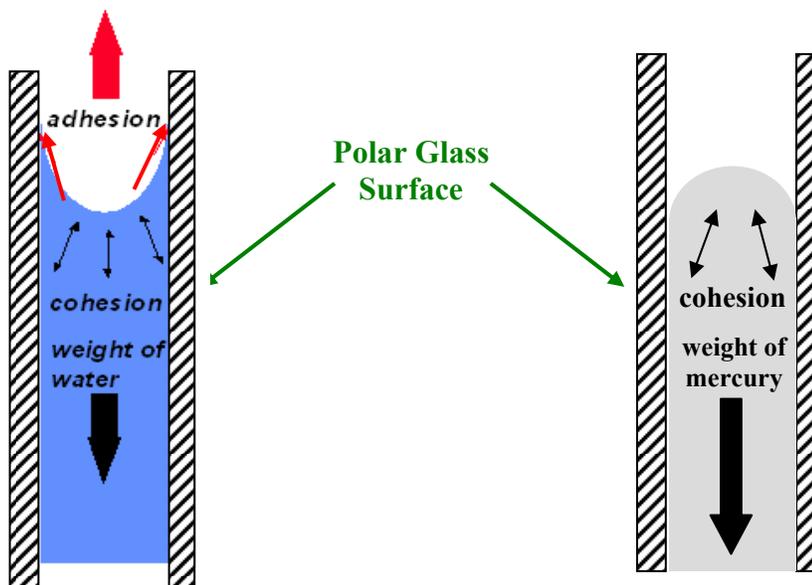


Surfactant: - a surface-active agent that decreases the surface tension. (Example: detergent)

Capillary Action: - when liquid suddenly rises in a narrow tube.

- there are two forces that causes capillary action to occur. They are cohesive forces and adhesive forces.

- Cohesive Forces:** - forces between molecules (intermolecular forces like London Dispersion Force, Dipole interactions, and Hydrogen Bonding).
- Adhesive Forces:** - forces between the liquid and the inner surface of the container.
 - in the case of capillary action, it would be the inside of the narrow tube.
 - polar liquid molecules with inner surface of the container that are also polar will create a large adhesive force (example: water and glass).
 - non-polar liquid molecules with polar inner surface of the container will create a small adhesive force (example: mercury and glass).



Both adhesion and cohesion forces are both strong.

However, adhesion > cohesion.

Therefore, water meniscus is concave.

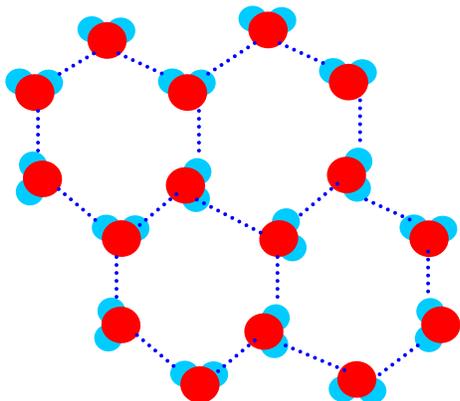
Cohesion forces are weak (Hg is non-polar).

There is no adhesion.

However, cohesion > adhesion.

Therefore, mercury meniscus is convex.

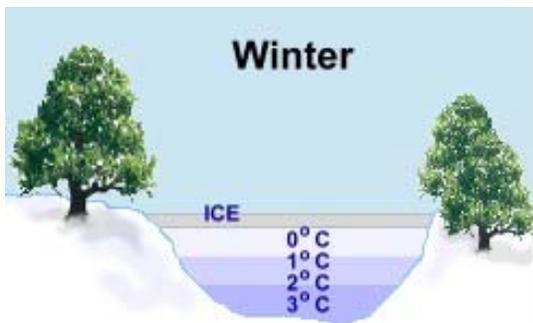
Hexagonal Arrangement of Ice Crystals: - due to the hydrogen bonds in water, it forms a honeycomb shape and expands in volume when it crystallizes into ice.
 - water expands as it freezes.



Due to the hydrogen bonds in water, it forms a *honeycomb shape* and expands in volume when it crystallizes into ice

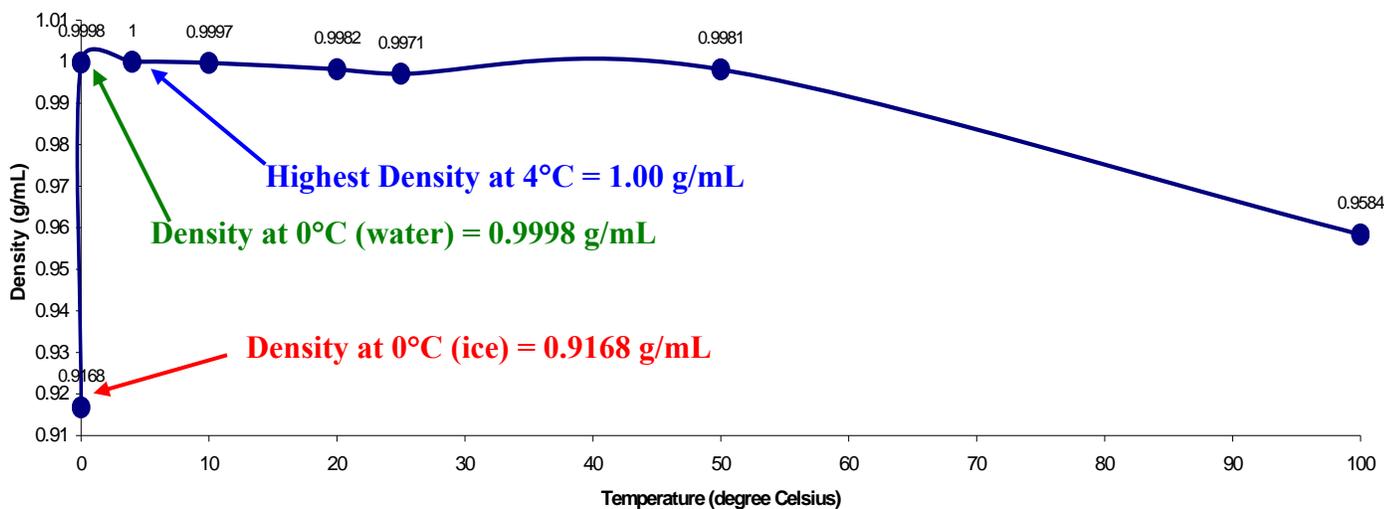
Even though no two snowflakes are alike, all of them have a basic hexagonal shape as dictated by the bent shape of water molecule and its hydrogen bonds

Density of Water: - water has the highest density at 4°C. Hence ice floats on water.



- because water expands when it freezes, it has a lower density in solid form. Therefore, ice floats on water.
- aquatic life can live in the cold climates because the warmer, denser water is at the bottom of the lake bed, and water freezes into ice from the top.

Density of Water



Assignment
 11.1 pg. 384 # 1 to 8, 11 to 13, 15
 11.2 pg. 384 # 1 to 8, 10 and 11

11.4: Phase Equilibrium

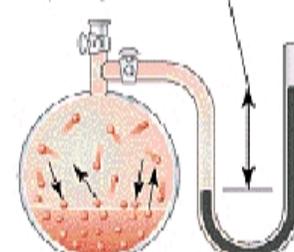
System: - a set components that are being studied.

Phase: - a region that has the same composition and properties throughout.

Dynamic Equilibrium: - when particles are constantly moving back and forth between phases in a system.

Vapour Pressure: - the pressure existed above a liquid when its rate of evaporation is the same as the rate of its condensation.
- sometimes refers to as **equilibrium vapour pressure** because **equilibrium** means the same rate of a two-way process.

$P_{\text{gas}} = \text{Equilibrium vapor pressure}$



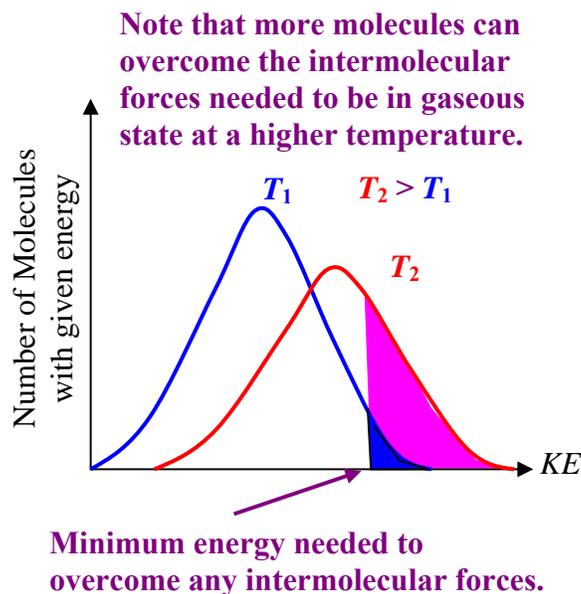
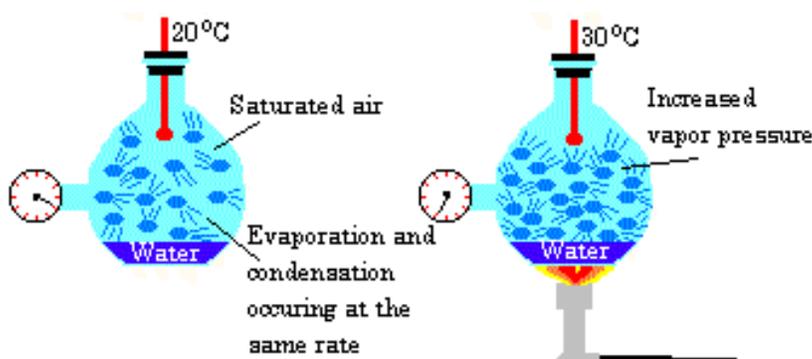
Vapour Pressure

$$P_{\text{atmosphere}} = P_{\text{vapour}} + P_{\text{Hg column}}$$

$$P_{\text{vapour}} = P_{\text{atmosphere}} - P_{\text{Hg column}}$$

- in general, for a particular compound, **as temperature increases, the vapour pressure increases.**
- This is because the **higher the temperature, kinetic energy increases** and **more liquid molecules will overcome the intermolecular forces** to become a gas. Thereby, **increasing vapour pressure.**

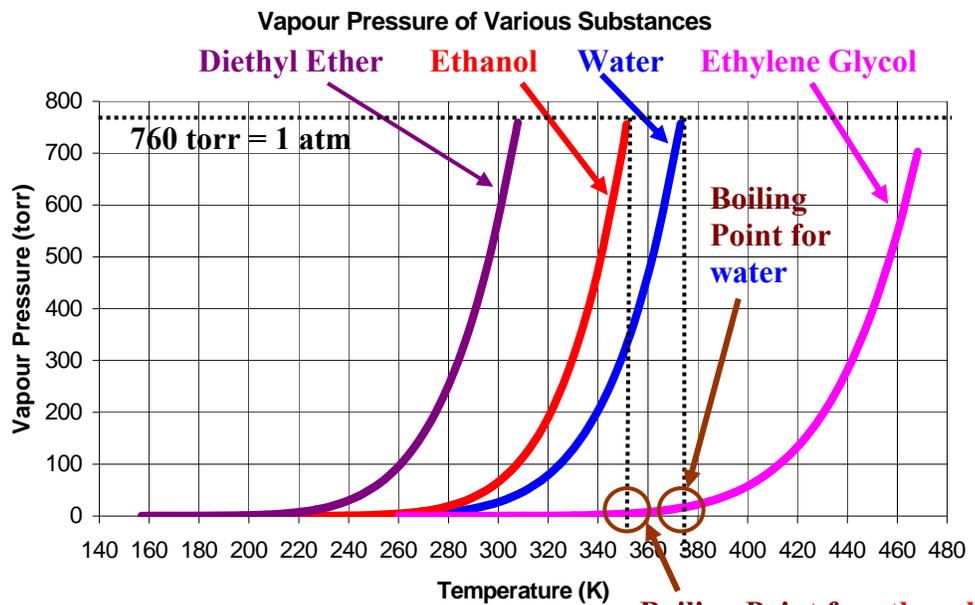
Temperature \uparrow , Vapour Pressure \uparrow



Volatile: - when liquids has a high vapour pressure.

- when comparing compounds at the same temperature, the compound that has a lower boiling point is considered more volatile because it will have a higher vapour pressure.
- when **intermolecular forces are weak**, the **compound is more volatile**. This is simply because it takes **less energy for the liquid to evaporate.**

We can use this graph to find the boiling points of various substances by moving across from a desired vapour pressure (which is equal to external air pressure at boiling). The temperature corresponds to that particular pressure for a certain substance is its boiling point at the air pressure.



Note: For substances, when $P_{\text{vap}} = 1 \text{ atm} = 760 \text{ torr}$, the corresponding temperature represents the boiling point.

Normal Melting Point: - the temperature where vapour pressures of liquid is equaled to that of a solid under normal atmospheric condition of 1 atm.

- when temperature is at the normal melting point, both liquid and solid coexist.
- when $T < \text{Melting Point}$, $P_{\text{vap}} \text{ Solid} < P_{\text{vap}} \text{ Liquid}$. Hence, any liquid's vapour pressure will be used by the solid to equilibrate its own vapour pressure. Thus, liquid will slowly become a solid below its melting point.
- when $T > \text{Melting Point}$, $P_{\text{vap}} \text{ Liquid} < P_{\text{vap}} \text{ Solid}$. As such, any solid's vapour pressure will be used by the liquid to equilibrate its own vapour pressure. Therefore, solid will slowly become a liquid above its melting point.

Normal Boiling Point: - the temperature where vapour pressures of liquid is equaled to that of a gas under normal atmospheric condition of 1 atm.

- when temperature is at the normal boiling point, both liquid and gas can coexist.
- when $T < \text{Boiling Point}$, $P_{\text{vap}} \text{ Liquid} < P_{\text{vap}} \text{ Gas}$. Hence, any vapour pressure of a gas will be used by the liquid to equilibrate its own vapour pressure. Thus, the gas will slowly become a solid below its boiling point.
- when $T > \text{Boiling Point}$, $P_{\text{vap}} \text{ Gas} < P_{\text{vap}} \text{ Liquid}$. As such, any liquid's vapour pressure will be used by the gas to equilibrate its own vapour pressure. Therefore, liquid will slowly become a gas above its boiling point.

Sublimation Point: - the temperature where vapour pressures of solid is equaled to that of a gas.

Normal Sublimation Point: - the temperature where vapour pressures of solid is equaled to that of a gas under standard atmospheric pressure of 1 atm.

Supercooled Liquid: - when a liquid is cooled too rapidly, its temperature is below the melting point.
- happens because molecules need time rearrange themselves for recrystallization to become a solid.

Superheated Liquid: - when a liquid is heated too rapidly, its temperature is above the boiling point.
- happens because molecules need time to completely break apart their intermolecular bonds to become a gas.
- the **bubbles of a superheated liquid tend to be large that burst violently** (known as **bumping**) as it's vapour pressure of the liquid is greater than that of the atmospheric pressure.

Phase Diagram: - a diagram with axes of Pressure versus Temperature to illustrate all three phases of a pure substance.
- it is more useful than a heating curve because boiling and melting points are pressure dependent, and the user now can see where the melting and boiling point are at any pressure. The curve of the vapour pressure versus temperature of a liquid is only a line on the phase diagram.
- it allows the user to identify the phase of a substance at a certain pressure and temperature.

Solid-Liquid Line: - a line that outlines the temperatures and the corresponding pressures where solid and liquid phases coexist in **dynamic equilibrium** (a state where the rate of forward process is exactly balanced by the reverse process).
- used to find the **melting point** of a substance at any given pressure.

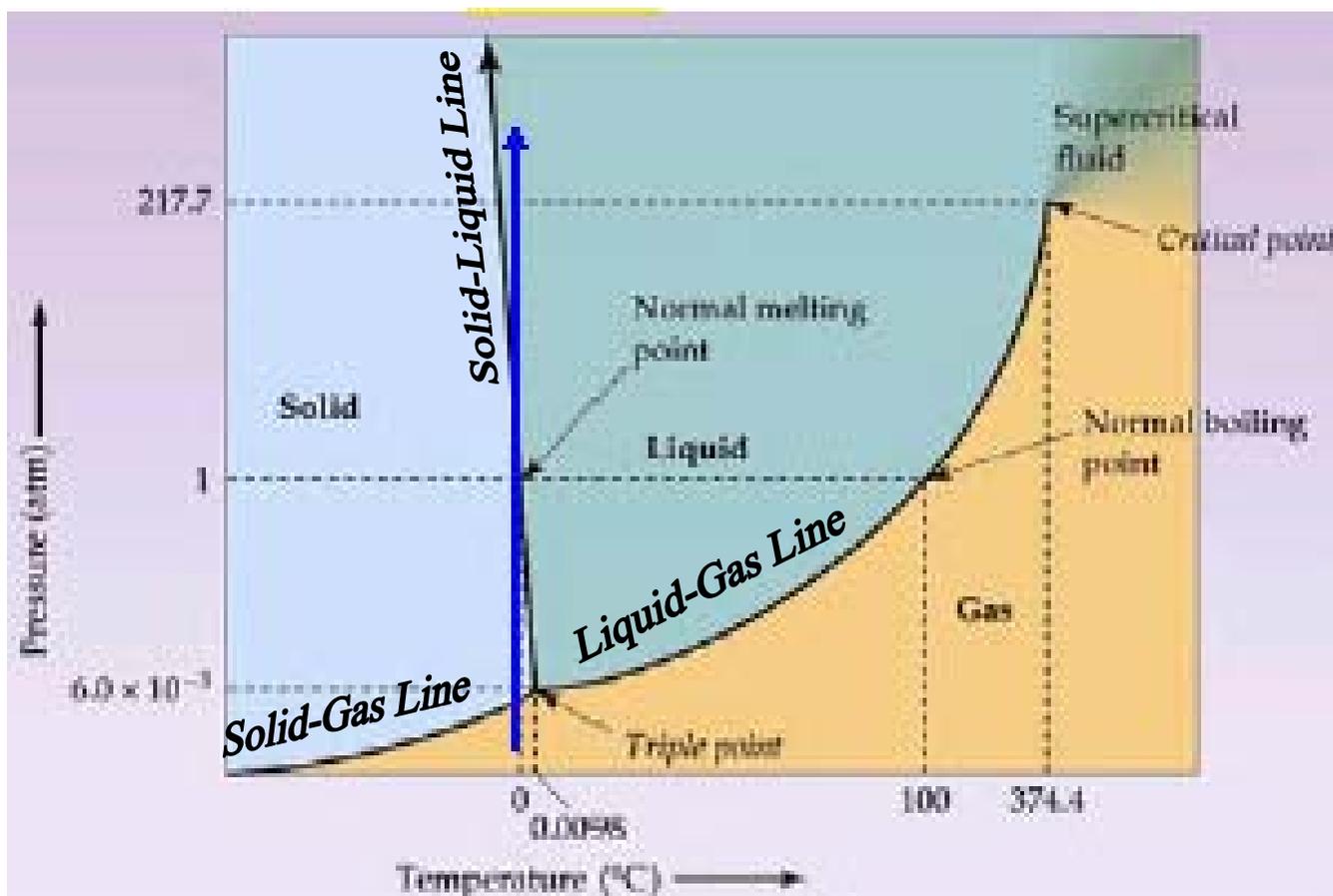
Liquid-Gas Line: - a line that outlines the temperatures and the corresponding pressures where liquid and gas phases coexist in dynamic equilibrium.
- used to find the **boiling point** of a substance at any given pressure.

Solid-Gas Line: - a line that outlines the temperatures and the corresponding pressures where solid and gas phases coexist in dynamic equilibrium.
- used to find the **sublimation point** of a substance at any given pressure.

Triple Point: - a point indicating the pressure and temperature where all solid-liquid, liquid-gas, and solid-gas lines meet.
- this is the **pressure and temperature conditions where all three phases can coexist in dynamic equilibrium.**

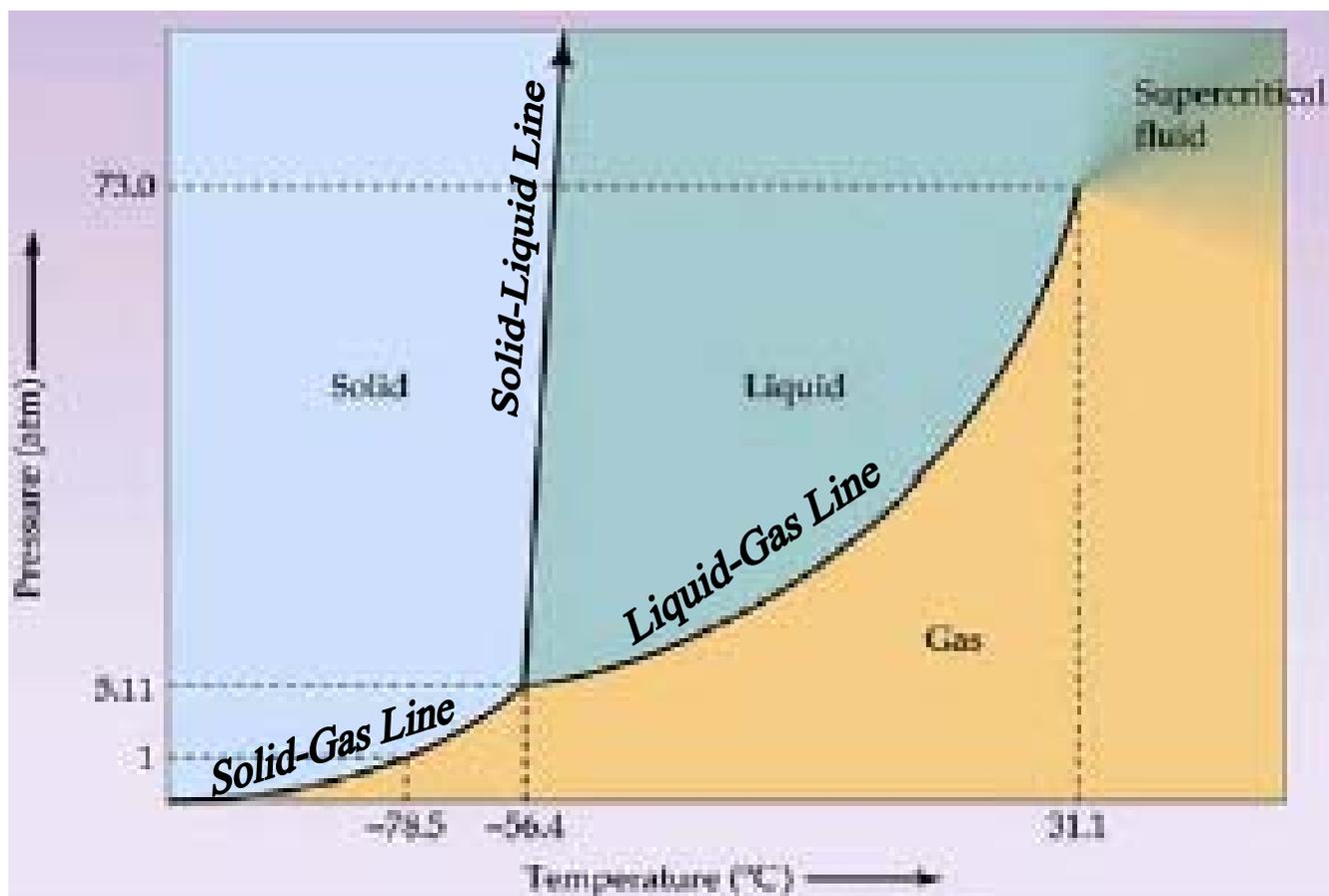
Critical Point: - a point at the end of the liquid-gas line which indicates an intermediate "fluid" region where liquid can be converted to gas instantaneously without waiting for phase change.
- the temperature at critical point is called **critical temperature** and the pressure at critical point is called **critical pressure.**
- at this region (beyond the critical pressure and temperature), the liquid will become a gas without going through a change of state.

Phase Diagram of Water



Notes from the Phase Diagram of Water:

- at 1 atm, water has a normal melting point of 0°C and a normal boiling point at 100°C.
- there is no normal sublimation point because the solid-gas line ends well under 1 atm (101.325 kPa).
- the negative slope of the solid-liquid line indicates that as pressure increases, the melting point of water decreases.
- the positive slope of the liquid-gas line illustrates that as pressure increases, the boiling point of water also increases.
- the positive slope of the solid-gas line shows that as pressure increases, the sublimation point of water increases as well.
- the triple point, where all three phases can coexist, is at 0.0098°C and 6.0×10^{-3} atm (0.608 kPa).
- the critical point of water is at 374.4°C and 217.7 atm (22 thousands kPa). Beyond this critical temperature and pressure, water will become gaseous instantaneously.

Phase Diagram of Carbon DioxideNotes from the Phase Diagram of Carbon Dioxide:

- at 1 atm (101.325 kPa), carbon dioxide has a normal sublimation point of -78.5°C .
- there is no normal melting and boiling points because carbon dioxide's solid-gas line is well above normal atmospheric condition of 1 atm.
- the positive slope of the solid-liquid line indicates that as pressure increases, the melting point of carbon dioxide increases.
- the positive slope of the liquid-gas line illustrates that as pressure increases, the boiling point of carbon dioxide also increases.
- the positive slope of the solid-gas line shows that as pressure increases, the sublimation point of carbon dioxide increases as well.
- the triple point, where all three phases can coexist, is at -56.4°C and 5.11 atm (518 kPa).
- the critical point of carbon dioxide is at 31.1°C and 73 atm (7400 kPa). Beyond this critical temperature and pressure, liquid carbon dioxide will become gaseous instantaneously.

Assignment

11.4 pg. 404 (Practice) #1; pg. 405 #1 to 10

Chapter 11 Review pg. 408–411 #9 to 20; 31 to 36, 45 to 47, 53, 60 to 65