

$$1 \text{ atm} = 14.7 \text{ psi}$$

Example

- A high-performance bicycle tire has a pressure of 132 psi. What is the pressure in mmHg?

$$132 \text{ psi} * \frac{1 \text{ atm}}{14.7 \text{ psi}} * \frac{760 \text{ mmHg}}{1 \text{ atm}}$$

$$= 6820 \text{ mmHg}$$

Gas Law Variables

- P = Pressure (atm)
- V = Volume (L)
- T = Temperature (K always)
- n = number of moles (moles)
- R = Gas Constant (L atm / mol K)

* Good Chart

Gas Laws

	Bolye's	Charles'	Avogrado's
Looked at	V,P	V,T	V,n
Held Constant	T,n	P,n	T,P
Found	Increase P Decrease V	Increase T Increase V	Increase n Increase V
Law	$P_1V_1 = P_2V_2$	$V_1/T_1 = V_2/T_2$	$V_1/n_1 = V_2/n_2$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$PV = nRT$$

$$\text{Example } \frac{782 \text{ torr} \times \frac{1 \text{ ATM}}{760 \text{ torr}}}{0.500 \text{ ATM}} = 1.03 \text{ ATM} = P_1$$

- A balloon is put in a bell jar and the pressure is reduced from 782 torr to 0.500 atm. If the volume of the balloon is now 2780 mL, what was it originally?

$$2780 = V_2$$

$$P_1 V_1 = P_2 V_2$$

$$V_1 = \frac{P_2 V_2}{P_1} = \frac{0.500 \times 2780 \text{ mL}}{1.03 \text{ ATM}}$$

$$V_1 = 1350 \text{ mL}$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{Example}$$

- The temperature inside a balloon is raised from 25.0°C to 250.0°C. If the volume of cold air was 10.0 L, what is the volume of hot air?

$$\frac{10.0 \text{ L} + 250.0^\circ\text{C}}{25.0^\circ\text{C}} = V_2 = 100 \text{ L}$$

$$\frac{10.0 \text{ L} + 250.0^\circ\text{C}}{298.2 \text{ K}} = V_2 = 17.5 \text{ L}$$

$$0^\circ\text{C} = 273.15 \text{ K}$$

$$25.0^\circ\text{C} + 273.15 \text{ K} =$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{Example}$$

- A 0.225 mol sample of He has a volume of 4.65 L. How many moles must be added to give 6.48 L?

$$\frac{4.65 \text{ L}}{0.225 \text{ mol}} = \frac{6.48 \text{ L}}{n_2} \quad n_2 = 0.314 \text{ mol}$$

The Ideal Gas Law

- By combining the gas laws we can write a general equation
- R** is called the **gas constant**

$$\underline{PV = nRT}$$

$$R = 0.08206 \frac{\text{L atm}}{\text{mol K}}$$

$$R = 0.08206 \frac{\text{L atm}}{\text{mol K}}$$

Example

- How many moles of gas are in a basketball with total pressure 24.3 psi, volume of 3.24 L at 25°C?

$$P = 24.3 \text{ Psi} \times \frac{1 \text{ atm}}{14.7 \text{ Psi}} = \boxed{1.65 \text{ atm}}$$

$$T = 25^\circ\text{C} + 273 \text{ K} = \boxed{298 \text{ K}}$$

$$n = \frac{PV}{RT} = \frac{1.65 \text{ atm} \times 3.24 \text{ L}}{0.08206 \frac{\text{L atm}}{\text{mol K}} \times 298 \text{ K}}$$

* $\boxed{n = 0.219 \text{ mol}}$

Standard Conditions

- Since the volume of a gas varies with pressure and temperature, chemists have agreed on a set of conditions to report our measurements so that comparison is easy – we call these **standard conditions**
 - STP
- Standard pressure = 1 atm
- Standard temperature = 273 K
– 0°C

Quiz: $n = ?$ when $V = 5.5 \text{ L}$, $P = 1.15 \text{ atm}$
 $T = 22.0^\circ\text{C}$

✓ $n = 0.26 \text{ mol}$

$$44.1 \text{ psi} \times \frac{1 \text{ atm}}{14.7 \text{ psi}} = 3 \text{ atm}$$

Example

$$T_1 = 300$$

- A gas occupies 10.0 L at 44.1 psi and 27°C. What volume will it occupy at standard conditions?

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$$

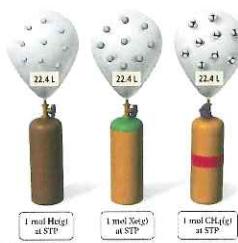
$$\underline{\text{STP} = 1 \text{ ATM}, 273.15 \text{ K}}$$

Know This

Molar Volume

- Solving the ideal gas equation for the volume of 1 mol of gas at STP gives 22.4 L
 - 6.022×10^{23} molecules of gas
 - notice: the gas is immaterial
- We call the volume of 1 mole of gas at STP the **molar volume**

Molar Volume



So $\frac{22.4 \text{ L}}{\text{mol}}$

as conversion factor

Quiz: Identify the gas

$$\text{mass} = 3.50 \text{ g}$$

$$V = 1.5 \text{ L}$$

$$P = 2.42 \text{ atm}$$

$$T = 355 \text{ K}$$

$$d = \frac{\rho}{L} = \frac{3.50 \text{ g}}{1.5 \text{ L}} = \boxed{2.33 \text{ g/L}}$$

$$d = \frac{mm \cdot P}{R \cdot T} \quad mm = \frac{d \cdot R \cdot T}{P} = \frac{(2.33 \text{ g/L}) (.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}) (355 \text{ K})}{(2.42 \text{ atm})}$$

~~mm = 28.048 g/mol~~ = ~~N₂~~

Density

- Density is the ratio of mass-to-volume
- Density of a gas is given in g/L
- The mass of 1 mole = molar mass
- The volume of 1 mole at STP = 22.4 L
- Density is directly proportional to molar mass at the same pressure and temperature
- Formula: $d = (P \times MM) / (R \times T)$

$$d = \frac{\text{molar mass} (P)}{R \cdot T}$$

Example

- Calculate the density of N₂ at 125°C and 755 mmHg

$$d = \frac{(0.993 \text{ atm})(28.02 \text{ L/mol})}{(0.08206 \frac{\text{L atm}}{\text{mol K}})(398 \text{ K})}$$

$$d = 0.852 \text{ g/L}$$

Thursday

Example

- Calculate the density of a gas at 775 torr and 27°C if 0.250 moles weighs 9.988 g

Quiz: Which will have the greatest Volume @ STP?



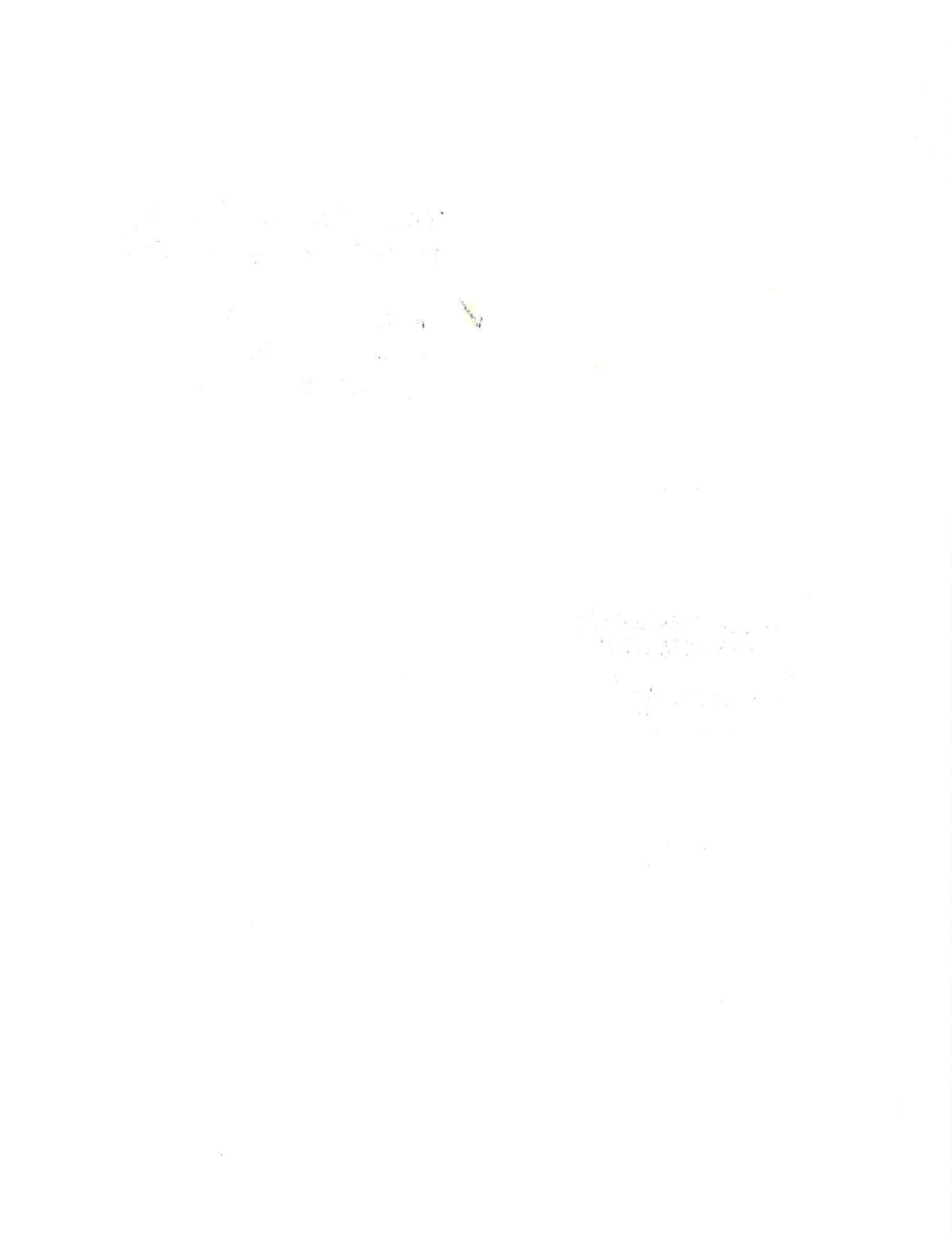
5g H₂

5g N₂

5g O₂

5g CO

All of the above



Mixing Gases

- When gases are mixed together, their molecules behave independent of each other
- Therefore, in certain applications, the mixture can be thought of as one gas

Partial Pressure

- The pressure of a single gas in a mixture of gases is called its **partial pressure**
- We can calculate the partial pressure of a gas if
 - we know what fraction of the mixture it composes and the total pressure
 - or, we know the number of moles of the gas in a container of known volume and temperature
- The sum of the partial pressures of all the gases in the mixture equals the total pressure
 - Dalton's Law of Partial Pressures

Avogadro Gas Law
states as # of moles
increase, so does Volume

Ex: filling up a balloon

Air

TABLE 5.3 Composition of Dry Air

Gas	Percent by Volume (%)
Nitrogen (N_2)	78
Oxygen (O_2)	21
Argon (Ar)	0.9
Carbon dioxide (CO_2)	0.04

not on test

$$P_{\text{total}} = P_A + P_B + P_C + \dots$$

$$P_A = P_T \cdot X_A \leftarrow \text{molar fraction}$$

$\downarrow O_2$ $\downarrow 1.00 \text{ atm}$ $\rightarrow 0.21$

$$P_A = P_T \cdot X_A \quad P_T = 1.00 \text{ atm}$$

Example $X_A = 0.009$

- Determine the pressure of Ar in the air at STP

$$(1.00 \text{ atm})(0.009) = 0.009 \text{ atm}$$

Mole Fraction

- the fraction of the total pressure that a single gas contributes is equal to the fraction of the total number of moles that a single gas contributes
- the ratio of the moles of a single component to the total number of moles in the mixture is called the **mole fraction, χ**

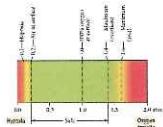
$$\chi_A = \frac{n_A}{n_{\text{total}}}$$

$$P_A = \chi_A \cdot P_{\text{total}}$$

$$\text{Ex: } O_2 = \frac{\text{mol } O_2}{\text{mol Air}}$$

Mountain Climbing & Partial Pressure

- Low = 0.21
High = 1.4
- Our bodies are adapted to breathe O_2 at a partial pressure of 0.21 atm
 - Sherpa, people native to the Himalaya mountains, are adapted to the much lower partial pressure of oxygen in their air
 - Partial pressures of O_2 lower than 0.1 atm will lead to **hypoxia**
 - unconsciousness or death
 - Climbers of Mt Everest carry O_2 in cylinders to prevent hypoxia



Go over This

$$P_T = P_{Ne} + P_{Xe} \quad PV = nRT$$

Example $n = \frac{PV}{RT}$

- Find the partial pressure of neon in a mixture with total pressure 3.9 atm, volume 8.7 L, temperature 598 K, and 0.17 moles Xe.

$$n = \frac{(3.9 \text{ atm})(8.7 \text{ L})}{(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(598 \text{ K})} = [0.69 \text{ mol}]$$

$$\text{mol Ne} = .69 - .17 = [0.52 \text{ mol Ne}]$$

$$X_{Ne} = \frac{0.52 \text{ mol}}{0.69 \text{ mol}} = [0.75]$$

$$\text{Partial Pressure} = 0.75 * 3.9 \text{ atm} = [2.9 \text{ atm}]$$

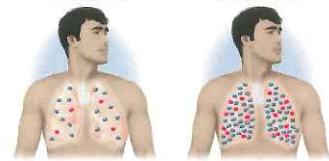
Deep Sea Divers and Partial Pressure

- It's also possible to have too much O₂, a condition called **oxygen toxicity**
 - P_{O₂} > 1.4 atm
 - oxygen toxicity can lead to muscle spasms, tunnel vision, and convulsions
- It's also possible to have too much N₂, a condition called **nitrogen narcosis**
 - also known as Rapture of the Deep
- When diving deep, the pressure of the air divers breathe increases – so the partial pressure of the oxygen increases

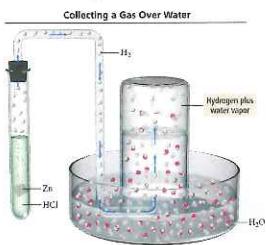
Partial Pressure & Diving

Surface
 $P_{N_2} = 1 \text{ atm}$
 $P_{O_2} = 0.78 \text{ atm}$
 $P_{CO_2} = 0.21 \text{ atm}$

30 m
 $P_{N_2} = 4 \text{ atm}$
 $P_{O_2} = 3.12 \text{ atm}$
 $P_{CO_2} = 0.84 \text{ atm}$



Collecting Gas by Water Displacement



Collecting Gases

- Gases are often collected by having them displace water from a container
- The problem is that since water evaporates, there is also water vapor in the collected gas
- The partial pressure of the water vapor, called the **vapor pressure**, depends only on the temperature
 - so you can use a table to find out the partial pressure of the water vapor in the gas you collect
 - Table 5.4*

Pg 201

TABLE 5.4 Vapor Pressure of Water versus Temperature			
Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
0	4.58	55	118.2
5	6.54	60	149.6
10	9.21	65	187.5
15	12.79	70	233.7
20	17.55	75	289.1
25	23.78	80	355.1
30	31.86	85	433.6
35	42.23	90	525.8
40	55.40	95	633.9
45	71.97	100	760.0
50	92.6		

not
on
Test

$$293\text{ K} = 20^\circ\text{C}$$

$$PV = nRT$$

$$P_{O_2} = 755.2 - P_{H_2O}$$

$$\text{Example } P_{H_2O} = 17.55$$

- 1.02 L of O₂ collected over water at 293 K with a total pressure of 755.2 mmHg. Find mass O₂.

$$n = \frac{PV}{RT} = \frac{(0.971)(1.02\text{ L})}{(0.08206)(293)}$$

$$P_{O_2} = 755.2 - 17.55$$

$$n = 0.0412 \text{ mol O}_2 * \frac{32.0\text{ g}}{\text{mol}}$$

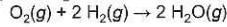
$$O_2 = 1.32\text{ g}$$

Reactions Involving Gases

- The principles of reaction stoichiometry from Chapter 4 can be combined with the gas laws for reactions involving gases
- In reactions of gases, the amount of a gas is often given as a volume
- The ideal gas law allows us to convert from the volume of the gas to moles; then we can use the coefficients in the equation as a mole ratio

Example @ STP
22.4 = 1 mol

- How many grams of H₂O form when 1.24 L H₂ reacts completely with O₂ at STP?



$$@STP = 1.24\text{ L H}_2 * \frac{2\text{ L H}_2O}{1\text{ L O}_2} = 2.48\text{ L H}_2O$$

$$2.48\text{ L H}_2O * \frac{1\text{ mol}}{22.4\text{ L}} * \frac{18.02\text{ g H}_2O}{\text{mol}} = 1.995$$

$$2.00\text{ g H}_2O$$

Properties of Gases

- Expand to completely fill their container
- Take the shape of their container
- Low density
- Compressible
- Mixtures of gases are always homogeneous
- Fluid

Kinetic Molecular Theory

- The particles of the gas (either atoms or molecules) are constantly moving
- The attraction between particles is negligible
- When the moving particles hit another particle or the container, they do not stick; but they bounce off and continue moving in another direction



Kinetic Molecular Theory

- There is a lot of empty space between the particles
- The average kinetic energy of the particles is directly proportional to the Kelvin temperature



Molecular Velocities

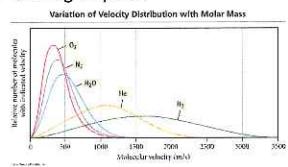
- All the gas molecules in a sample can travel at different speeds
- However, the distribution of speeds follows a pattern called a **Boltzman distribution**
- The “average velocity” of the molecules, but there are different ways to take this kind of average
- The method of choice for our average velocity is called the **root-mean-square** method.

Kinetic Energy and Molecular Velocities

- Average kinetic energy of the gas molecules depends on the average mass and velocity
– $KE = \frac{1}{2}mv^2$
- Gases in the same container have the same temperature, the same average kinetic energy
- If they have different masses, the only way for them to have the same kinetic energy is to have different average velocities

Molecular Speed vs. Molar Mass

- in order to have the same average kinetic energy, heavier molecules must have a slower average speed



Temperature and Molecular Velocities

- $KE_{avg} = \frac{1}{2}N_A m u^2$
- N_A is Avogadro's number
- $KE_{avg} = 1.5RT$
- R is the gas constant in energy units, $8.314 \text{ J/mol}\cdot\text{K}$
- $1 \text{ J} = 1 \text{ kg}\cdot\text{m}^2/\text{s}^2$
- Equating and solving we get:
- $N_A \cdot \text{mass} = \text{molar mass in kg/mol}$

$$u_{rms} = \sqrt{\frac{3RT}{N_A \cdot m}} = \sqrt{\frac{3RT}{MM}}$$

u_{rms} = Speed of gas Particle

$$\text{Speed} = \sqrt{\frac{3RT}{MM}}$$

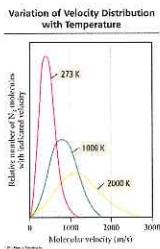
$$\text{Molar Mass} = \frac{\text{kg}}{\text{mol}}$$

$$\sqrt{\frac{(3/\cancel{\text{mol}})(\cancel{K})}{\text{kg/mol}}} = \sqrt{\frac{J}{\text{kg}}} = \sqrt{\frac{\text{kg}\cdot\text{m}^2}{\text{s}^2}} \frac{\text{kg}}{\text{s}}$$

$$\sqrt{\frac{\text{m}^2}{\text{s}^2}} = \boxed{\frac{\text{m}}{\text{s}}}$$

Temperature vs. Molecular Speed

- As the absolute temperature increases, the average velocity increases



$$O_2 = \frac{32.0 \cancel{\text{g}}}{\text{mol}} * \frac{1 \cancel{\text{kg}}}{1000 \cancel{\text{g}}} = \frac{0.03200 \text{ kg}}{\text{mol}}$$

- Calculate the rms velocity of O_2 at 25°C

$$v_{rms} = \sqrt{\frac{3(8.314 \cancel{\text{J/mol}}\text{K})(298 \text{K})}{0.03200 \text{ kg/mol}}}$$

$$v_{rms} = 481.946 \approx \boxed{482 \text{ m/s}}$$

Effusion Example:

$\frac{\text{He}}{\text{O}_2}$

$$\sqrt{\frac{32.00 \text{ g/mol}}{4.00 \text{ g/mol}}} = 2.83 \times \text{faster}$$

Put the bigger
molecules on top

Diffusion and Effusion

- The process of a collection of molecules spreading out from high concentration to low concentration is called **diffusion**
- The process by which a collection of molecules escapes through a small hole into a vacuum is called **effusion**
- Both the rates of diffusion and effusion of a gas are related to its rms average velocity

Quiz: Which will travel faster at STP?

Ne

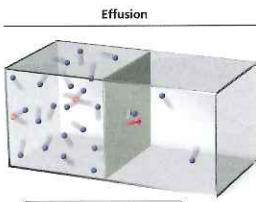
Xe

✓ He

CO

CO₂

Effusion



Gas escapes from container into a vacuum through a small hole

Diffusion = The bigger the molecule the slower it takes to saturate a room

Effusion = Gases only from an area to a vacuum

Graham's Law of Effusion

- For two different gases at the same temperature, the ratio of their rates of effusion is given by the following equation:

$$\frac{\text{rate}_{\text{gas A}}}{\text{rate}_{\text{gas B}}} = \sqrt{\frac{\text{Molar Mass}_{\text{gas B}}}{\text{Molar Mass}_{\text{gas A}}}}$$

Example

- Calculate the molar mass of a gas that effuses at a rate 0.462 times N_2

$$\sqrt{0.462 N_2} = \sqrt{\frac{28.02 \text{ g/mol}}{?}}$$

$$0.213 = \frac{28.02 \text{ g/mol}}{X}$$

$$X = \frac{28.02}{0.213} = 131.3 \text{ g/mol}$$

Number given is greater than One 1
then unknown gas is on Top

Number given is less than One 1
Unknown gas is on bottom

~~Ideal vs. Real Gases~~

- Real gases often do not behave like ideal gases at high pressure or low temperature
- Ideal gas laws assume
 - There are no interactions
 - Based on the kinetic-molecular theory
- At low temperatures and high pressures these assumptions are not valid

The End

Chapter 6

"Thermochemistry"

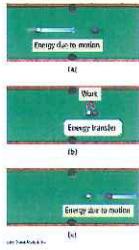


Nature of Energy

- Even though Chemistry is the study of matter, energy effects matter
- **Energy** is anything that has the capacity to do work
- **Work** is a force acting over a distance
 - Energy = Work = Force x Distance
- Energy can be exchanged between objects through contact

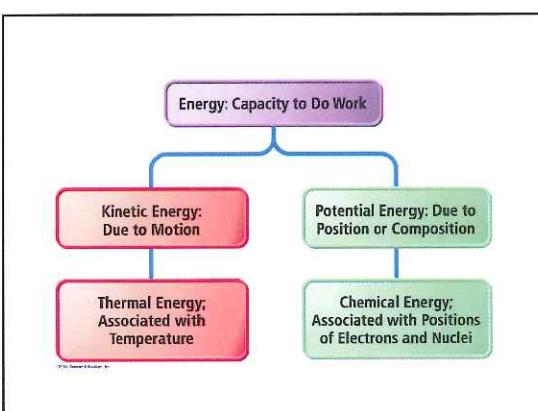
Classification of Energy

- **Kinetic energy** is energy of motion or energy that is being transferred
 - thermal energy is kinetic



Classification of Energy

- Potential energy** is energy that is stored in an object, or energy associated with the composition and position of the object
 - energy stored in the structure of a compound is potential



Potential = Position

Law of Conservation of Energy

- Recall: energy cannot be created or destroyed
 - First Law of Thermodynamics
- Energy can be transferred between objects
- Energy can be transformed from one form to another
 - heat → light → sound

The diagram shows two people. One person is pushing a ball on a horizontal surface, with a label "Mechanical potential energy" pointing to the ball. The other person is pulling on a spring, with a label "Kinetic energy" pointing to the spring. This illustrates how energy can be transferred between objects and transformed from one form to another.

Some Forms of Energy

- Electrical — *electrons*
– kinetic energy
- Heat or Thermal Energy ~ *Atoms, molecules*
– kinetic energy
- Light or Radiant Energy ~ *Photons*
– kinetic energy
- Nuclear — *Neutrons to Protons*
– potential energy
- Chemical — *mainly electron*
– potential energy

Units of Energy

- The amount of kinetic energy an object has is directly proportional to its mass and velocity
✓ $KE = \frac{1}{2}mv^2$
- When the mass is in kg and speed in m/s, the unit for kinetic energy is
- 1 joule of energy is the amount of energy needed to move a 1 kg mass at a speed of 1 m/s
✓ $1 J = 1 \frac{\text{kg} \cdot \text{m}^2}{\text{s}^2}$

$KE = \text{Kinetic Energy}$

$KE = \frac{1}{2} m \cdot v^2$

$m = \text{kg}$

$V = \text{m/s}$

Units of Energy

- **joule** (J) is the amount of energy needed to move a 1 kg mass a distance of 1 meter
– $1 J = 1 \text{ N} \cdot \text{m} = 1 \text{ kg} \cdot \text{m}^2/\text{s}^2$
- **calorie** (cal) is the amount of energy needed to raise one gram of water by 1°C
– kcal = energy needed to raise 1000 g of water 1°C
– food Calories = kcals
– $1 \text{ cal} = 4.184 \text{ J}$ on HW

$1 \text{ Cal} = \text{Food Calorie}$

$1 \text{ Kcal} = 1 \text{ Cal}$

Energy Use				
Unit	Energy Required to Raise Temperature of 1 g of Water by 1°C	Energy Required to Light 100-W Bulb for 1 hr	Energy used to Run 1 Mile (approx)	Energy Used by Average U.S. Citizen in 1 Day
joule (J)	4.18	3.60×10^5	4.2×10^5	9.0×10^8
calorie (cal)	1.00	8.60×10^4	1.0×10^5	2.2×10^8
Calorie (Cal)	0.00100	86.0	100.	2.2×10^5
kWh	1.16×10^{-6}	0.100	0.12	2.5×10^2

Energy Flow and Conservation of Energy

- we define the **system** as the material or process we are studying the energy changes within
- we define the **surroundings** as everything else in the universe
- Conservation of Energy requires that the total energy change in the system and the surrounding must be zero
 - $\Delta\text{Energy}_{\text{universe}} = 0 = \Delta\text{Energy}_{\text{system}} + \Delta\text{Energy}_{\text{surroundings}}$
 - Δ is the symbol that is used to mean change
 - final amount - initial amount

Need to Know

Internal Energy

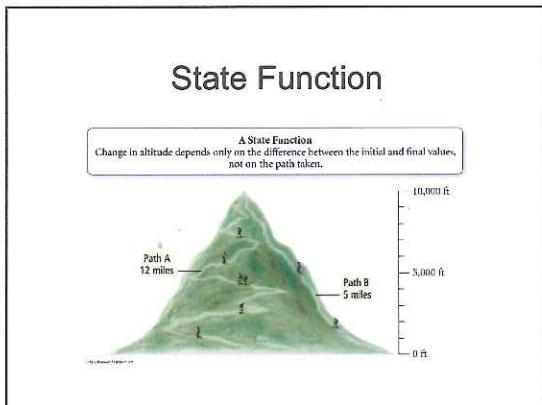
- the **internal energy** is the total amount of kinetic and potential energy a system possesses
- the change in the internal energy of a system only depends on the amount of energy in the system at the beginning and end
 - a **state function** is a mathematical function whose result only depends on the initial and final conditions, not on the process used
 - $\Delta E = E_{\text{final}} - E_{\text{initial}}$
 - $\Delta E_{\text{reaction}} = E_{\text{products}} - E_{\text{reactants}}$

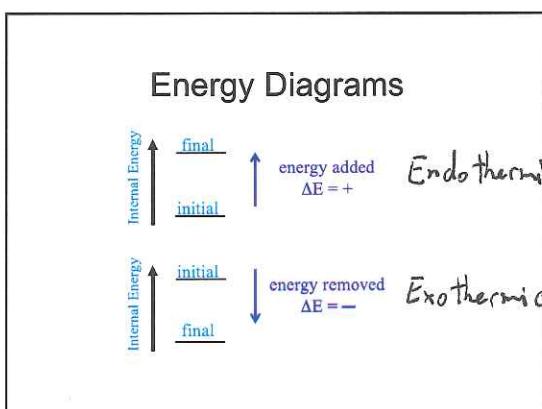
System is the chemical reaction

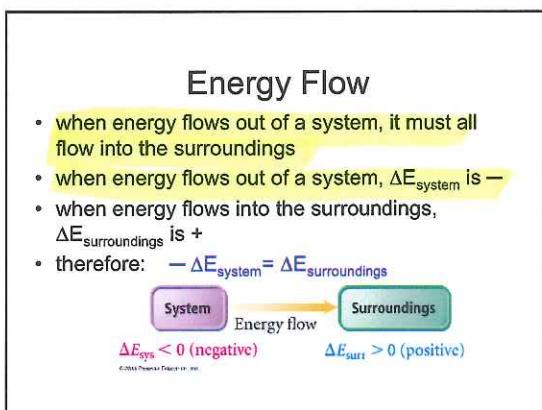
Surrounding is everything else

E = energy Internal

Is the system giving energy or the surrounding?







System gives energy
→ Surrounding takes it

Surrounding gives energy
→ System gains it

Energy out = Energy in

Energy Flow

- when energy flows into a system, it must all come from the surroundings
- when energy flows into a system, ΔE_{system} is +
- when energy flows out of the surroundings, $\Delta E_{\text{surroundings}}$ is -
- therefore: $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$



How Is Energy Exchanged?

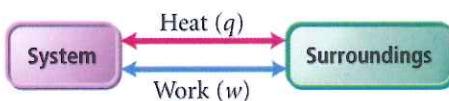
- energy is exchanged between the system and surroundings through heat and work
 - q = heat (thermal) energy
 - w = work energy
 - q and w are NOT state functions, their value depends on the process

$$\Delta E = q + w$$

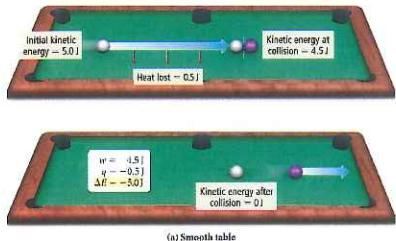
q (heat)	system gains heat energy +	system releases heat energy -
w (work)	system gains energy from work +	system releases energy by doing work -
ΔE	system gains energy +	system releases energy -

Work = like when gas expands it is doing work + losing energy

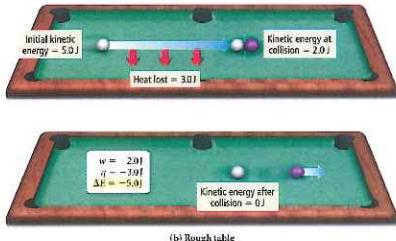
Energy Exchange



Heat & Work (smooth table)



Heat & Work (rough table)



Heat Exchange

- Heat is the exchange of thermal energy between the system and surroundings
- Heat flows from matter with high temperature to matter with low temperature until both objects reach the same temperature
 - thermal equilibrium

Quantity of Heat Energy Absorbed Heat Capacity

- When a system absorbs heat, its temperature increases the increase in temperature is directly proportional to the amount of heat absorbed
- The proportionality constant is called the **heat capacity, C**
— units of C are J/C or J/K
- $q = C \times \Delta T$
- The heat capacity of an object depends on its mass
- The heat capacity of an object depends on the type of material

Quiz: $m = 50\text{ g}$, $q = 3456\text{ J}$

$$C_{\text{water}} = 4.18 \frac{\text{J}}{\text{g}^\circ\text{C}} \quad \Delta T = ?$$

$$\Delta T = \frac{q}{m \cdot C}$$

$$\Delta T =$$

Specific Heat Capacity

- Measure of a substance's **intrinsic** ability to absorb heat
- The **specific heat capacity** is the amount of heat energy required to raise the temperature of one gram of a substance 1°C
— units are $\text{J}/(\text{g}^\circ\text{C})$
- The **molar heat capacity** is the amount of heat energy required to raise the temperature of one mole of a substance 1°C

Substance	Specific Heat Capacity, C_s $\text{J}/(\text{g}^\circ\text{C})$
Elements	
Lead	0.126
Gold	0.123
Silver	0.125
Copper	0.205
Iron	0.449
Aluminum	0.903
Compounds	
Ethanol	2.42
Water	4.18
Metals	
Chromium	0.39
Mercury	0.33
Sand	0.64

The higher
the number
the more
heat it
can take
in

* pg 241

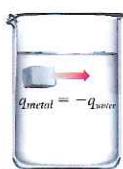
$\Delta T = T_{\text{final}} - T_{\text{initial}}$

Specific Heat -Calculation

Heat = (mass) x (specific heat capacity) x (temp. change)

$$q = (m) \times (C_s) \times (\Delta T)$$

q \downarrow
 m \downarrow
 C_s \downarrow
 ΔT \downarrow



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

Example

- How much heat is absorbed by a copper penny with mass 3.10 g whose temperature rises from -8.0°C to 37.0°C ?

$$q = (3.10\text{g})(45.0^\circ\text{C})(0.385 \frac{\text{J}}{\text{g}^\circ\text{C}})$$

$$q = 53.7 \text{ J}$$

Pressure -Volume Work

- PV work is work that is the result of a volume change against an external pressure
- When gases expand, ΔV is +, but the system is doing work on the surroundings so w is -
- As long as the external pressure is kept constant
Work = External Pressure x Change in Volume
 $w = -P\Delta V$

Note: to convert the units to joules use $101.3 \text{ J} = 1 \text{ atm} \cdot \text{L}$

Pg 24S

$$101.3 \text{ J} = 1 \text{ atm} \cdot \text{L}$$

$$\text{Conversion factor} = \frac{101.3 \text{ J}}{1 \text{ atm} \cdot \text{L}}$$

$$\Delta V = 1.75 \text{ L}$$

Example

- If a balloon is inflated from 0.100 L to 1.85 L against an external pressure of 1.00 atm , how much work is done?

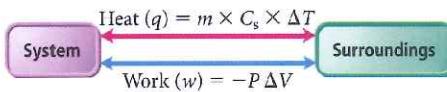
$$1.75 \text{ L} \cdot 1.00 \text{ atm} = 1.75 \text{ L} \cdot \text{atm}$$

$$w = -1.75 \text{ L} \cdot \text{atm} * \frac{101.3 \text{ J}}{1 \text{ atm} \cdot \text{L}}$$

$$= -177 \text{ J}$$

$$w = -P\Delta V$$

Exchanging Energy Between System and Surroundings



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Measuring ΔE , Calorimetry at Constant Volume

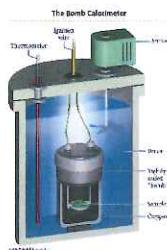
- $\Delta E = q + w$, therefore we can determine ΔE by measuring q and w
- It is easiest to do a process in such a way that there is no change in volume, $w = 0$
- It is not possible to observe the temperature changes of the individual chemicals involved in a reaction – so we use an insulated, controlled surroundings and measure the temperature change in it

$$q_{\text{surroundings}} = q_{\text{calorimeter}} = -q_{\text{system}}$$

$$-\Delta E_{\text{reaction}} = q_{\text{cal}} = C_{\text{cal}} \times \Delta T$$

\downarrow \rightarrow
 5°C 0°C

Bomb Calorimeter



Example

- When 1.010 g of sugar is burned in a bomb calorimeter, the temperature rises from 24.92°C to 28.33°C. If $C_{cal} = 4.90 \text{ kJ}/\text{C}$, find ΔE for burning 1 mole

$$1.010 \text{ g} * \frac{1}{342.32 \text{ J/mol}} = .00295 \text{ mol}$$

$$q_{surround} = \frac{4.90 \text{ kJ}}{\text{C}} * 3.35 \text{ C} = 16.7 \text{ kJ}$$

$$q_{system} = -16.7 \text{ kJ}$$

$$\Delta E = -16.7 \text{ kJ} = \boxed{-5660 \text{ kJ/mol}}$$

Sucrose = $C_{12}H_{22}O_{11} = 342.32 \text{ J/mol}$

* We can never calculate the system, we can only calculate the surrounding & reverse it for the system

Enthalpy

- The **enthalpy**, **H**, of a system is the sum of the internal energy of the system and the product of pressure and volume
 $- H$ is a state function $H = E + PV$
- The **enthalpy change**, ΔH , of a reaction is the heat evolved in a reaction at constant pressure $\Delta H_{reaction} = q_{reaction \text{ at constant pressure}}$
- Usually ΔH and ΔE are similar in value, the difference is largest for reactions that produce or use large quantities of gas

4 different possible ways of calculating ΔH

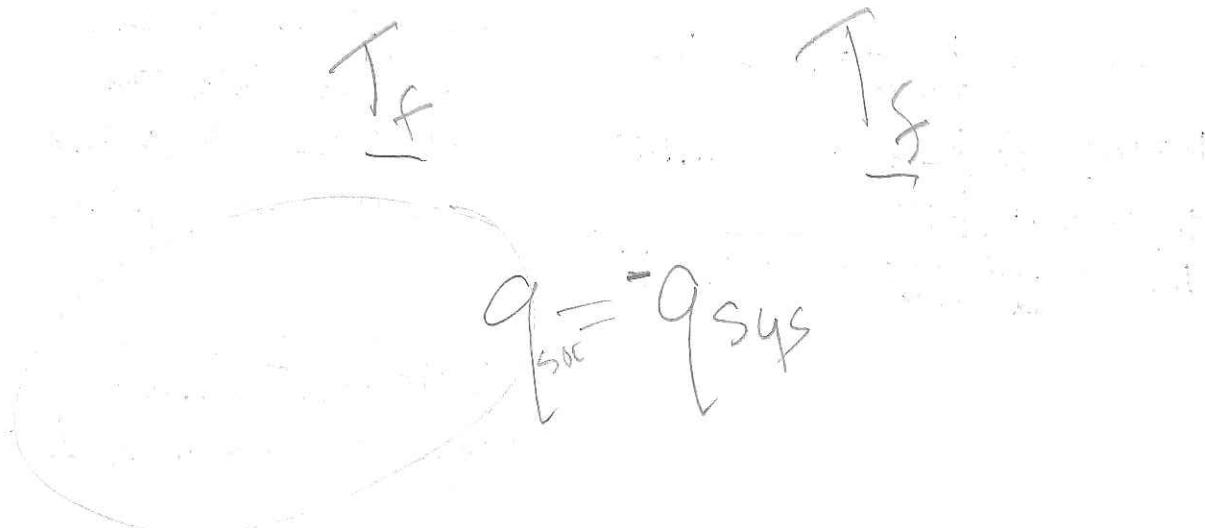
Endothermic and Exothermic Reactions

- When ΔH is $-$, heat is being released by the system
- Reactions that release heat are called **exothermic reactions**
- When ΔH is $+$, heat is being absorbed by the system
- Reactions that absorb heat are called **endothermic reactions**

absorb



$$\Delta T = T_f - T_i$$



$$Gold = m = 31.7 \text{ g}$$

$$T_i = 69.6^\circ\text{C}$$

$$c = 0.128 \text{ J/g°C}$$

$$Water = m = 63.8 \text{ g}$$

$$T_i = 26.5^\circ\text{C}$$

$$c = 4.18 \text{ J/g°C}$$

Heat gained by water = Heat lost by gold

$$\text{heat lost by gold} = q = m \cdot c \cdot \Delta T$$

$$q_{gold} = 31.7 \text{ g} * 0.128 \text{ J/g°C} * (69.6 - T_f)$$

$$\text{heat lost by water} = q = m \cdot c \cdot \Delta T$$

$$q_{water} = 63.8 \text{ g} * 4.18 \text{ J/g°C} * (T_f - 26.5)$$

$$31.7 \text{ g} * 0.128 \text{ J/g°C} * (69.6^\circ\text{C} - T_f) = 63.8 \text{ g} * 4.18 \text{ J/g°C} * (T_f - 26.5^\circ\text{C})$$

$$4.05 \text{ J/g°C} (69.6^\circ\text{C} - T_f) = 266.7 \text{ J/g°C} (T_f - 26.5^\circ\text{C})$$

$$281.8 \text{ J} - 4.05 T_f = 266.7 T_f - 7014.2 \text{ J}$$

$$270.75 T_f = 7296 \text{ J}$$

$$\boxed{T_f = 26.9^\circ\text{C}}$$

Quiz: How much heat is evolved from 229g of C₃H₈

$$229 \text{ g} * \frac{\text{mol}}{44.11 \text{ g/mol}} * \frac{-2044 \text{ kJ}}{\text{mol C}_3\text{H}_8} = -10600 \text{ kJ}$$

Enthalpy of Reaction

- The enthalpy change in a chemical reaction is an extensive property
 - the more reactants you use, the larger the enthalpy change
 - By convention, we calculate the enthalpy change for the number of moles of reactants in the reaction as written
- C₃H₈(g) + 5 O₂(g) → 3 CO₂(g) + 4 H₂O(g) ΔH = -2044 kJ

- ΔH_{reaction} for 1 mol C₃H₈ = -2044 kJ
 - ΔH_{reaction} for 5 mol O₂ = -2044 kJ
- depends which product we use

$$13.2 \text{ kg} * \frac{1000 \text{ g}}{\text{kg}} * \frac{\text{mol}}{44.11 \text{ g}} = 299.25 \text{ mol}$$

Example

- How much heat is evolved in the complete combustion of 13.2 kg of C₃H₈(g)?

$$299.25 \text{ mol} * \frac{-2044 \text{ kJ}}{1 \text{ mol C}_3\text{H}_8}$$

$$= -6.12 \times 10^5 \text{ kJ}$$

$$1 \text{ mol C}_3\text{H}_8 = -2044 \text{ kJ}$$

Measuring ΔH Calorimetry at Constant Pressure

- Reactions done in aqueous solution are at constant pressure
 - open to the atmosphere
- The calorimeter is often nested foam cups containing the solution

$$q_{\text{reaction}} = -q_{\text{solution}} = -(m_{\text{solution}} \times C_{\text{s, solution}} \times \Delta T)$$

- ΔH_{reaction} = q_{constant pressure} = q_{reaction}
 - to get ΔH_{reaction} per mol, divide by the number of moles



Lecture 32:30

Know This Calculation

10/18/10

$$q = m \cdot c \cdot \Delta T$$

$$q = 100.2 \text{ g} \cdot 4.18 \frac{\text{J}}{\text{g}\cdot\text{C}} \cdot 7.2^\circ\text{C}$$

Example

- What is $\Delta H_{rxn}/\text{mol Mg}$ for the reaction $\text{Mg(s)} + 2 \text{HCl(aq)} \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})$
- If 0.158 g Mg reacts in 100.0 mL of solution changes the temperature from 25.6°C to 32.8°C?

$$q = 3.0 \text{ kJ surround} = -3.0 \text{ kJ}$$

$$\Delta H = \left(\frac{-3.0 \text{ kJ}}{0.158 \text{ g}} \right) * \left(\frac{24.3 \text{ g}}{\text{mol}} \right)$$

$$\Delta H = \boxed{-460 \text{ kJ/mol}}$$

Relationships Involving ΔH_{rxn}

- When reaction is multiplied by a factor, ΔH_{rxn} is multiplied by that factor
- $\text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \quad \Delta H = -393.5 \text{ kJ}$
- $\text{C(s)} + 2 \text{O}_2(\text{g}) \rightarrow 2 \text{CO}_2(\text{g}) \quad \Delta H = 2(-393.5 \text{ kJ}) = 787.0 \text{ kJ}$
- If a reaction is reversed, then the sign of ΔH is reversed
- $\text{CO}_2(\text{g}) \rightarrow \text{C(s)} + \text{O}_2(\text{g}) \quad \Delta H = +393.5 \text{ kJ}$

$$\Delta T = 7.2^\circ\text{C}$$

$$m = 100 \text{ mL water} * \text{density}$$

$$m = 100.0 \text{ mL} * 1.00 \frac{\text{g}}{\text{mL}} = 100.0 \text{ g H}_2\text{O}$$

$$m = 100.0 \text{ g H}_2\text{O} + 0.158 \text{ g Mg} = 100.2 \text{ g}$$

$$\Delta H = \frac{\text{kJ}}{\text{mol}}$$

Quiz: if the rxn in the calorimeter was endothermic, the Temp would?

increase

decrease

remain same

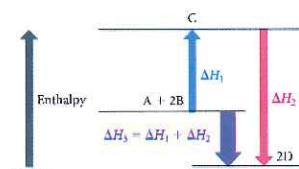
cannot determine

unknown

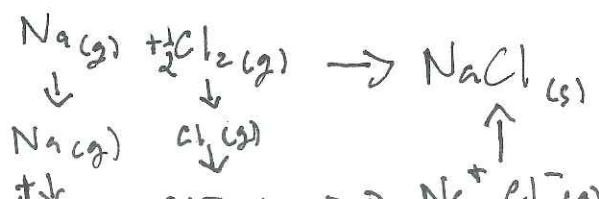
Relationships Involving ΔH_{rxn}

Hess's Law

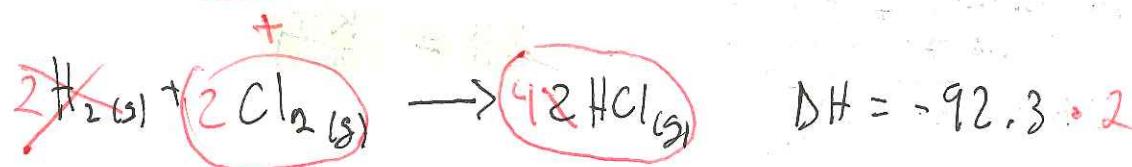
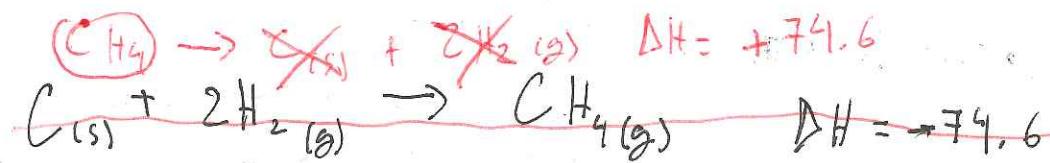
Hess's Law
The change in enthalpy for a stepwise process is the sum of the enthalpy changes of the steps.



Ex:

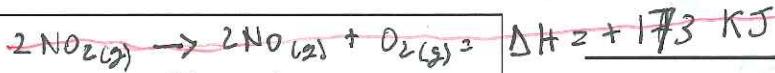
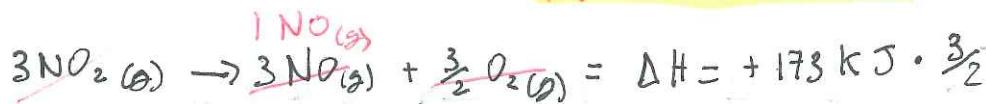


PG 272

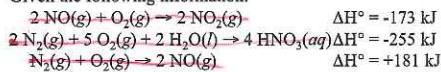
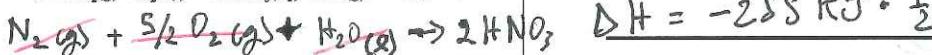
 $E_x = 80$ 

$$= 74.6 \text{ kJ} + (-95.7 \text{ kJ}) + (-92.3 \text{ kJ} \times 2)$$

$$= -205.7 \text{ kJ} = \boxed{-206 \text{ kJ}}$$

**Example**

Given the following information:

Calculate the ΔH° for the reaction below:

$$\Delta H = (173 \cdot \frac{3}{2}) + (-255 \cdot \frac{1}{2}) + (-181 \text{ kJ})$$

$$\boxed{\Delta H = -49 \text{ kJ}}$$

Standard Conditions

- The **standard state** is the state of a material at a defined set of conditions.
 - pure gas at exactly 1 atm pressure
 - pure solid or liquid in its most stable form at exactly 1 atm pressure and temperature of interest (usually 25°C)
 - substance in a solution with concentration 1 M
- The **standard enthalpy change**, ΔH° , is the enthalpy change when all reactants and products are in their standard states
- The **standard enthalpy of formation**, ΔH_f° , is the enthalpy change for the reaction forming 1 mole of a pure compound from its constituent elements
 - the elements must be in their standard states
 - ΔH_f° for a pure element in its standard state = 0 kJ/mol

1 atm, 25 °C, 1 M

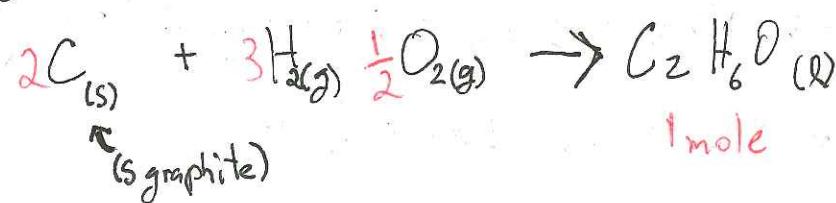
The "little "o" on ΔH
means standard conditions

Pg 272 #81 practice**Writing Formation Reactions**

- Write the formation reaction for $\text{CO}(g)$
- The formation reaction is the reaction between the elements in the compound
 $\text{C} + \text{O} \rightarrow \text{CO}(g)$
- The elements must be in their standard state
 $\text{C(s, graphite)} + \text{O}_2(g) \rightarrow \text{CO}(g)$
- The equation must be balanced, but the coefficient of the product compound must be 1
 $\text{C(s, graphite)} + \frac{1}{2} \text{O}_2(g) \rightarrow \text{CO}(g)$

Lecture ~20 min

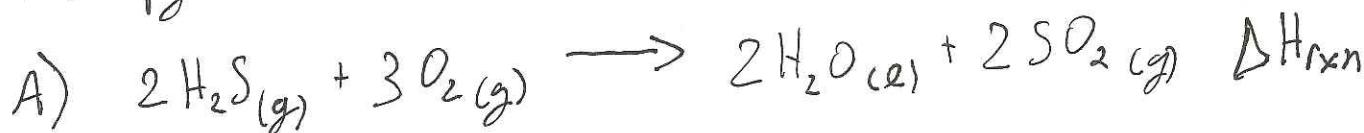
Ex: formation Equation



Exam Example
#86 pg 272

(products - reactant)

using ΔH_f°



Tables Appendix II

$$\Delta H_{rxn} \left[\left(2 \frac{\text{mol}}{\text{mol}} \cdot -285.8 \frac{\text{kJ}}{\text{mol}} \right) + \left(2 \text{mol} \cdot -296.8 \frac{\text{kJ}}{\text{mol}} \right) \right] =$$

-571.6 kJ products -593.6

$$\left[\left(2 \text{mol} \cdot -20.6 \frac{\text{kJ}}{\text{mol}} \right) + \left(0 \frac{\text{kJ}}{\text{mol}} \cdot 3 \text{mol} \right) \right]$$

reactant

$$\Delta H_{rxn} = (-1165.2 \text{ kJ}) - (-41.2 \text{ kJ}) = \boxed{-1124.0 \text{ kJ}}$$

* Sig figs - multiply first & keep sig figs truncated in rules of multiplication. Now add values & keep sig figs under rule of addition even if you end up with more digits to accommodate for addition law

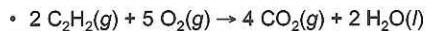
Practice #83 + 85

Calculating Standard Enthalpy Change for a Reaction

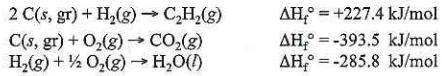
- Any reaction can be written as the sum of formation reactions
 - The ΔH° for the reaction is then the sum of the ΔH_f° for the component reactions
- $\Delta H^\circ_{\text{reaction}} = \sum n \Delta H_f^\circ(\text{products}) - \sum n \Delta H_f^\circ(\text{reactants})$ *Know This*
- Σ means sum
 - n is the coefficient of the reaction

Equation not given on Exam

Calculate the Enthalpy Change in the Reaction (1st way)



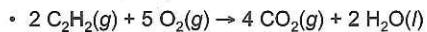
Write formation reactions for each compound and determine the ΔH_f° for each



Arrange equations so they add up to desired reaction

$$\begin{aligned} [(4 \cdot -393.5) + (2 \cdot -285.8)] - \\ [2 \cdot 227.4] &= -2600.6 \\ &= \boxed{-2601 \text{ kJ}} \end{aligned}$$

Calculate the Enthalpy Change in the Reaction (2nd way)



$$\Delta H^\circ_{\text{reaction}} = \sum n \Delta H_f^\circ(\text{products}) - \sum n \Delta H_f^\circ(\text{reactants})$$

$$\Delta H_{\text{rxn}} = [(4 \cdot \Delta H_{\text{CO}_2} + 2 \cdot \Delta H_{\text{H}_2\text{O}}) - (2 \cdot \Delta H_{\text{C}_2\text{H}_2} + 5 \cdot \Delta H_{\text{O}_2})]$$

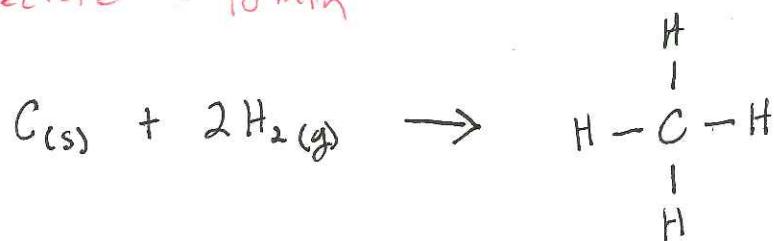
$$\Delta H_{\text{rxn}} = [(4 \cdot (-393.5) + 2 \cdot (-285.8)) - (2 \cdot (+227.4) + 5 \cdot (0))]$$

$$\Delta H_{\text{rxn}} = -2600.4 \text{ kJ}$$

$$\Delta H_{\text{rxn}}^\circ = \sum n \Delta H_f^\circ(\text{products}) - \sum n \Delta H_f^\circ(\text{reactants})$$

Lecture ~48 min

Ex:



Page 387
for chart

ΔH_{rxn} = bonds broken - bonds made

ΔH_{rxn} = Break (2 \cdot H-H bond) - (4 \cdot C-H bonds) make

$$\Delta H_{rxn} = (2 \cdot 436 \frac{kJ}{mol}) - (4 \cdot 414 \frac{kJ}{mol})$$

$$\Delta H_{rxn} = 872 \text{ kJ} - 1660 \text{ kJ} = -788 = \boxed{-790 \text{ kJ}}$$

sig figs to the tenth position

————— // ————— // —————

Determine ΔH_{rxn} for the following rxn using
bond dissociation energies

Exam!!



$$[(5 \cdot C-H \text{ bond}) + (C-C \text{ bond}) + (C-O \text{ bond}) + (O-H \text{ bond}) + (3 O=O)]$$

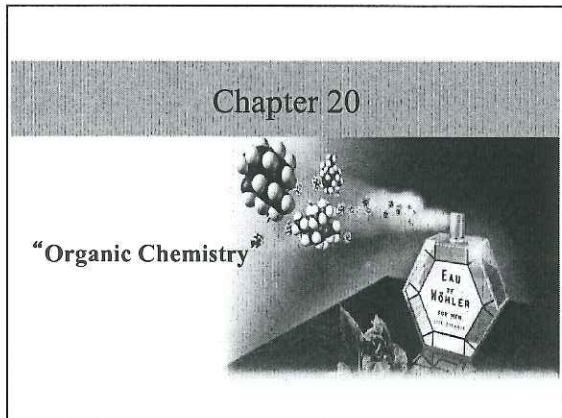
$$- [(4 \cdot C=O) + (6 \cdot O-H)]$$

$$\begin{aligned} * C=O \\ * CO_2 \text{ bond} = 799 \frac{kJ}{mol} \end{aligned}$$

$$= [(5 \cdot 414) + (347) + (360) + (464) + (3 \cdot 498)] - [(799 \cdot 4) + (6 \cdot 464)]$$

$$= 4735 - 5980 = \boxed{-1245 \text{ kJ}}$$

The End



Why Carbon?

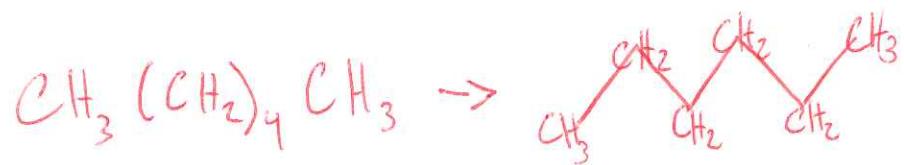
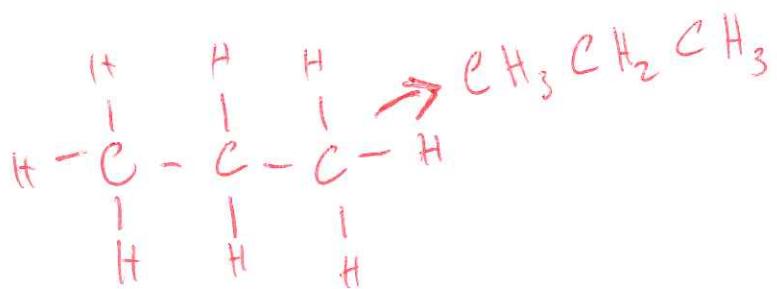
- Usually it forms four bonds
- It forms covalent bonds
- It forms polar covalent bonds
- It can form double and triple bonds
- The shape of organic molecules are predicted by VSEPR theory
- Can be sp^3 , sp^2 or sp

• it is very small

Small size = close proximity
= Strong bonds

Carbon and Hydrogen:

- Hydrocarbon
- Four main types
 - Alkanes: only single bonded carbons
 - Alkenes: contains at least one double bond
 - Alkynes: contains at least one triple bond
 - Aromatics: contain alternating double and single bonds in a ring with a pi electron cloud
- Isomers: same molecular formula but different structural formula



Hydrocarbons

- Saturated:
 - All single bonds
 - No more hydrogen can be added
- Unsaturated:
 - One or more double or triple bonds
 - Can add more hydrogen to the structure

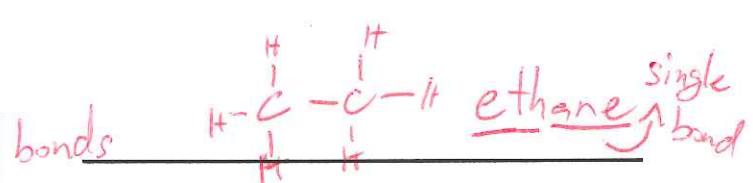
Nomenclature

- Prefix – express the number of carbons
- Root – expresses the carbon/carbon connectivity
- Suffix- tell the functional group

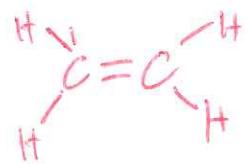
Prefix = How many Carbon bonds

- Prefixes
- 1 - Meth-
- 2 - Eth-
- 3 - Prop-
- 4 - But-
- 5 - Pent

- Prefixes
- 6 - Hex
- 7 - Hept
- 8 - Oct
- 9 - Non
- 10 - Dec



the last "e" in Ethane & Ethene = only hydrogen bonds

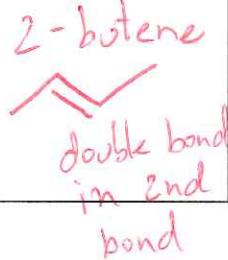
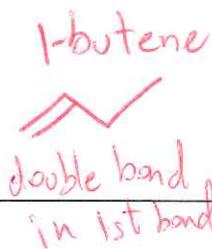


Ethene ^{double bond} 2

Nomenclature

- Root

- For all single bonds – an
- For at least one double bond – en
- For at least one triple bond – yn



Nomenclature

- Suffix

- Based on the functional group
- For all hydrocarbon -e

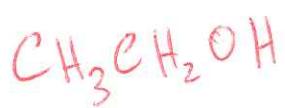
Ex: Ethane

Functional Groups

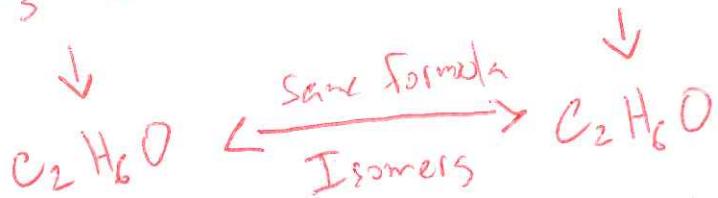
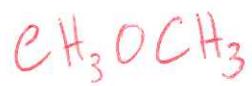
- Defined: part of the molecule in which the main chemistry takes place
- No functional group in alkanes
- The double bond is the functional group in alkene
- The triple bond is the functional group in alkynes

Functional Group is where the double bond (or Triple) is located. That is where the reaction takes place

Alcohol



Ether



Exam: Only name Hydrocarbon with e at the end

4/17/12

Functional Groups (Oxygen)

- Alcohol

- ROH

- Name ends with -ol



Methanol
single bond
C₂H₅O

- Ether

- ROR

- Name ends with -ate



Ethol
C₂H₅O

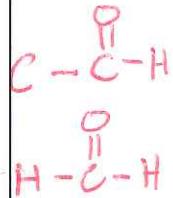
The "R" = any Carbon Group

Functional Groups (Oxygen)

- Aldehyde

- RCHO

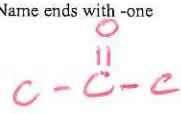
- Name ends with -al



- Ketone

- RCOR

- Name ends with -one

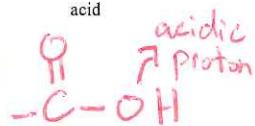


Functional Groups (Oxygen)

- Carboxylic Acid

- RCOOH

- Name ends with -oic and acid



- Ester *Smell good*

- RCOOR

- Name group with single oxygen first, then prefix for the other side then root -oate



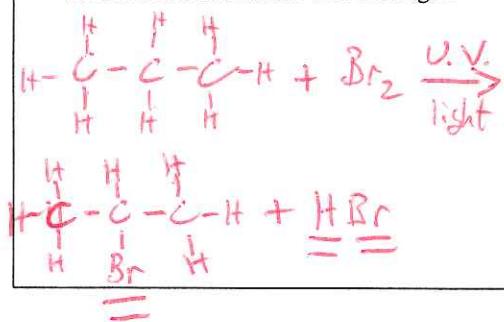
Functional Groups (Nitrogen)

Presence of N

- Amine
 - RNH_2 , R_2NH , R_3N
 - Name groups attached to the nitrogen then end with amine
- Amide
 - RCOCH_2 , RCONR , RCO_2
 - Name ends in amide

A Few Common Reactions

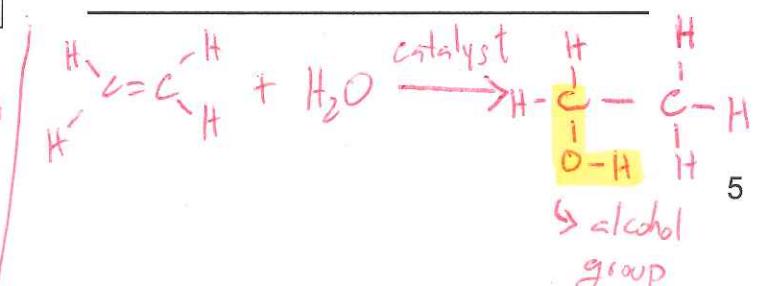
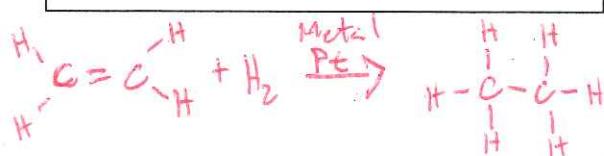
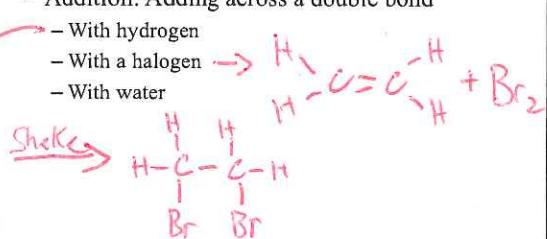
- Substitution: An alkane with a halogen



A Few Common Reactions

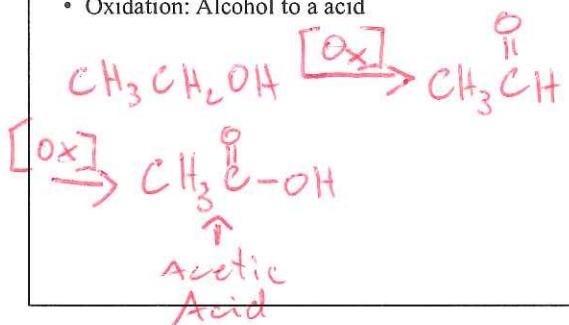
- Addition: Adding across a double bond

- With hydrogen
- With a halogen
- With water



A Few Common Reactions

- Oxidation: Alcohol to a acid



The End
