The Mole Guided Notes

What is the Mole?

- Just like 1 dozen equals 12, 1 mole equals ______________________________.
- It's sometimes called _____________________ number.
- 1 mole can be abbreviated as _______________

How large is it?

- 1 mole of _____________________ is $6,020,000,000,000,000,000,000.
- 1 mole of _____________________ would make a planet the same size as earth!
- 1 mole of _____________________ would cover the earth 5 miles (8 km) deep!

Why do we count in moles?

A teaspoon of water contains about 4.94 x 10^{23} water molecules.

- That's about 0.8 moles.
- Since there are a _____________________________ in a small amount of chemical, we count in terms of moles.

We don't actually count at all….

How many pennies are in a 50.0 lb bag?

- Instead of counting them, _____________ the bag and ______________ a single penny.
- Then, you can use ______________ to figure out the number of pennies in the bag!

Atomic Mass = Mass of 1 Mole!

- How much does 1 mole of carbon weigh?
- How much does 1 mole of calcium weigh?

1 atom of S = 32.06 amu

6.02 x 10^{23} atoms of S = 32.06 g

Atomic Mass = Molar Mass

- Atomic mass tells the...
  - Mass of 1 atom (__________)
  - Mass of 1 mole (__________)
- ALWAYS round to ______ decimal places

Practice Problems: Molar Mass of ELEMENTS

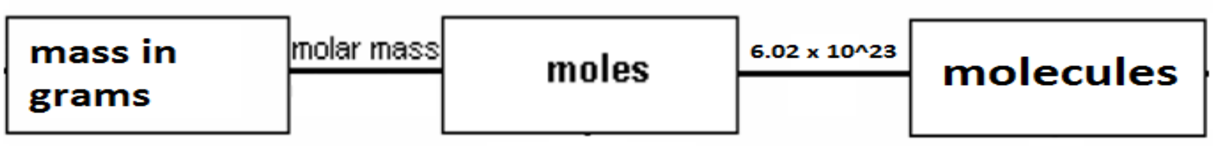
- magnesium ___________
- fluorine ___________
- sodium ___________

Practice Problems: Molar Mass of COMPOUNDS

- water ___________
- sodium chloride ___________
- calcium hydroxide ___________
- ammonium sulfide___________

Molar Conversions

Use the Factor-Label Method!



Practice Problem:

- How many moles of carbon are in 46 g of carbon?

Practice Problem:

- How many molecules are in 3.50 moles of $Ca(OH)_2$?

Practice Problem:

- Find the mass of 4.1×10^{24} molecules of $(NH_4)_2S$.

THE MOLE

WHAT IS THE MOLE?

- Just like 1 dozen equals 12, **1 mole equals 6.02×10^{23}.**
- It's sometimes called Avogadro's number.
- 1 mole can be abbreviated as **1 mol.**

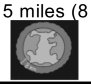
Amadeo Avogadro

HOW LARGE IS THE MOLE?

- 1 mole of pennies is $6,020,000,000,000,000,000,000.

- 1 mole of basketballs would make a planet the same size as earth!

- 1 mole of donuts would cover the earth 5 miles (8 km) deep!



WHY DO WE COUNT IN MOLES?

- A teaspoon of water contains about 4.94×10^{23} water molecules.
- That's about 0.8 moles.
- Since there are a LOT of molecules in a small amount of chemical, we count in terms of moles.

WE DON'T ACTUALLY COUNT AT ALL…

How many pennies are in a 50.0 lb bag?

- Instead of counting them, weigh the bag and weigh a single penny.
- Then, you can use MATH to figure out the number of pennies in the bag!

ATOMIC MASS = MASS OF 1 MOLE!

- How much does 1 mole of **carbon** weigh?

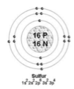
  12.01 grams

- How much does 1 mole of **calcium** weigh?

  40.08 grams

ATOMIC MASS = MOLAR MASS

- Atomic mass tells both the…
  - mass of 1 atom (in amu)
  - mass of 1 mole (g/mol)



Sulfur

1 atom of S = 32.06 amu

6.02×10^{23} atoms of S = 32.06 g

ALWAYS Round to 2 decimal places

PRACTICE PROBLEMS: MOLAR MASS OF ELEMENTS

- magnesium **24.31 g/mol**
- fluorine **19.00 g/mol**
- sodium **22.99 g/mol**

PRACTICE PROBLEMS: MOLAR MASS OF COMPOUNDS

- water
  - H_2O
  - 2(1.01) + 16.00 = **18.02 g/mol**
- sodium chloride
  - NaCl
  - 22.99 + 35.45 = **58.44 g/mol**

PRACTICE PROBLEMS: MOLAR MASS OF COMPOUNDS

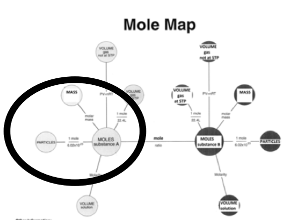
- calcium hydroxide
  - $Ca(OH)_2$
  - 40.08 + 2(16.00) + 2(1.01)

= 74.1 g/mol → **74.10 g/mol**

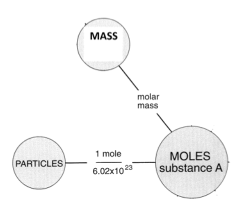
- ammonium sulfide
  - $(NH_4)_2S$
  - 2(14.01) + 8(1.01) +32.06

= **68.16 g/mol**

MOLAR CONVERSIONS



MOLAR CONVERSIONS

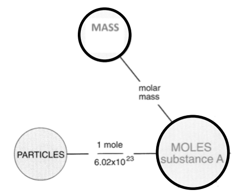


USE THE FACTOR-LABEL METHOD!

- How many moles of carbon are in 46 g of carbon?

46 g C	

=



USE THE FACTOR-LABEL METHOD!

- How many moles of carbon are in 46 g of carbon?

$$\frac{46\ \cancel{g}\ C}{} \left| \frac{1\ mol\ C}{12.01\ \cancel{g}\ C} \right. = \mathbf{3.8304...\ mol\ C}$$

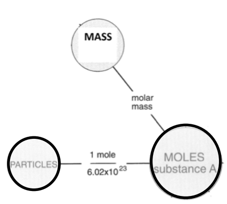
= **3.8304… mol C → 3.8 mol C**

PRACTICE PROBLEM:

- How many molecules are in 3.50 moles of $Ca(OH)_2$?

3.50 mol	

=



PRACTICE PROBLEM:

- How many molecules are in 3.50 moles of $Ca(OH)_2$?

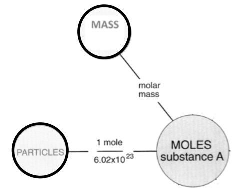
$$\frac{3.50\ \cancel{mol}}{} \left| \frac{6.02 \times 10^{23}\ molecules}{1\ \cancel{mol}} \right. = \mathbf{2.107 \times 10^{24}\ molecules\ Ca(OH)_2}$$

= 2.107×10^{24} → 2.11×10^{24} molecules $Ca(OH)_2$

PRACTICE PROBLEM:

- Find the mass of 4.1×10^{24} molecules of $(NH_4)_2S$.

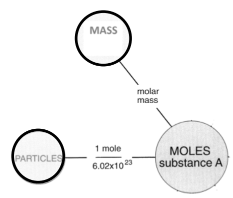
4.1×10^{24} molecules	



PRACTICE PROBLEM:

- Find the mass of 4.1×10^{24} molecules of $(NH_4)_2S$.

4.1×10^{24} ~~molecules~~	1 mol	
	6.02×10^{23} ~~molecules~~	



PRACTICE PROBLEM:

- Find the mass of 4.1×10^{24} molecules of $(NH_4)_2S$.

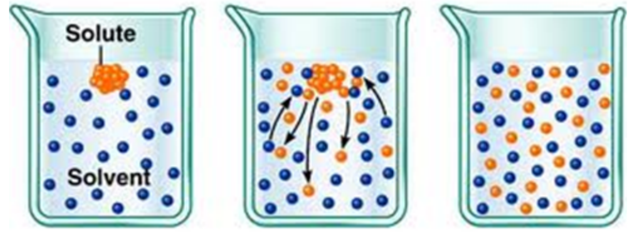
4.1×10^{24} ~~molecules~~	1 ~~mol~~	68.16 g	**= 464.2126… g $(NH_4)_2S$**
	6.02×10^{23} ~~molecules~~	1 ~~mol~~	

= 464.2126… → 460 g $(NH_4)_2S$

Solutions Notes

Lesson 1

A solution is a ______________________________ (usually involving water and another substance). The other substance can be a gas, solid, or another liquid. In a solution, the dissolved substance is ______________________ mixed with the water.



Homogeneous mixture:
Heterogeneous mixture:

Solute –
Solvent –

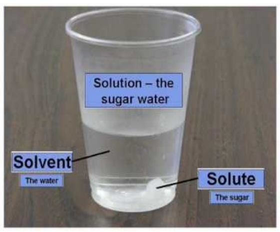


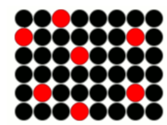
TABLE 14.1 Some Common Solutions

Solution	Components
Gaseous solutions	
Air	N_2, O_2, and several others
Natural gas	CH_4, C_2H_6, and several others
Liquid solutions	
Seawater	H_2O, NaCl, and many others
Vinegar mL of	H_2O, $HC_2H_3O_2$ (acetic acid)
Soda pop	H_2O, CO_2, $C_{12}H_{22}O_{11}$ (sucrose), and several others
Solid solutions	
Yellow brass	Cu, Zn
Palladium–hydrogen	Pd, H_2

Suspensions/Colloids: ______________________ mixtures; Suspensions settle out and colloids don't settle

Example suspensions – orange juice; paint; Italian dressing

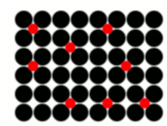
Example colloid – mayonnaise; fog; milk



Substitution Alloy

Tyndall Effect – is _______________________ scattered or not?

Alloys – mixtures of a metal and a nonmetal (_________________) or 2 metals (_________________).



Interstitial Alloy

Alloys – NOT _________________. The mixture does have unique properties, though.
Examples: Nitinol, Steel
Solutions have different Concentrations! Concentrations are measured in ______________________ (____).

$$6M = \frac{6\ mol}{1\ L}$$

<u>Practice Problem:</u>
How many grams of NaCl are needed to make 0.500 L of 0.25M NaCl?

Mole Fractions – a ratio of _________; it's always moles solute divided by total moles solution

$$X_A = \frac{\text{moles A}}{\text{Total moles}}$$

<u>Practice Problem:</u>
If 10.5g of sodium chloride and 12.6g of potassium chloride are dissolved in 250.g of water, what is the mole fraction of the chloride ion?

Practice Problem:

The density of acetonitrile, CH_3CN, is 0.786g/mL. and the density of methanol, CH_3OH, is 0.791g/mL. A solution is made by dissolving 20.0mL CH_3OH in 100.0mL of CH_3CN. The vapor pressure of pure CH_3CN is 81.0 mmHg at 22°C.

a. What is the mole fraction of methanol in the solution?

b. What is the new vapor pressure of the CH_3CN with the methanol in solution at 22°C? (Ignore methanol's vapor pressure).

c. Assuming that the volumes are additive, what is the molarity of the solution?

Practice Problem:

A solution was prepared by dissolving 23.7g of $CaCl_2$ in 375g of water. The density of the resulting solution was 1.05g/mL. What is the molarity of the Cl^- ion in this solution? Assume 100% dissociation of the compound. The molar mass of $CaCl_2$ is 110.984g/mol.

Lesson 2

Miscible - Liquids __________________ with other liquids

Immiscible - Liquids __________________ with other liquids

What is really happening when a solution is made?

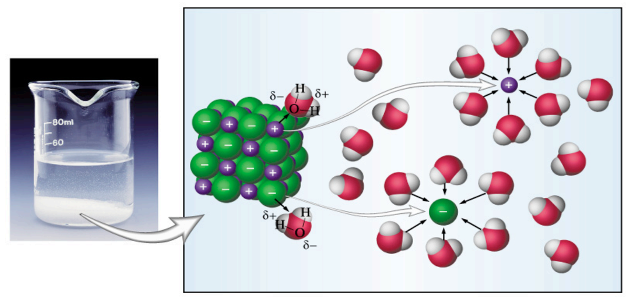
1) Solute particles must break apart (overcome intermolecular forces); this is ______________________ (requires energy)
2) Solvent particles must break apart (overcome intermolecular forces); this is ______________________ (requires energy)
3) Solute particles & solvent particles mix and form new intermolecular forces; this is ______________________ (releases energy)
4) The sum of these energy changes is known as enthalpy of ________________ (______)

Ideal Solution – dissolving causes __________ energy change $\Delta H_{soln} = 0$

Non-ideal Solutions:

1) $\Delta H_{soln} < 0$; _______thermic; intermolecular forces between solute-solvent are _____________ than the IMF within the solute and the IMF within the solvent.
2) $\Delta H_{soln} > 0$; _______thermic; intermolecular forces between solute-solvent are _____________ than the IMF within the solute and the IMF within the solvent; solution may not form.
3) $\Delta H_{soln} >> 0$; VERY _______thermic; intermolecular forces between solute-solvent are MUCH _____________ than the IMF within the solute and the IMF within the solvent; solution DOES NOT form…heterogen. mixture (suspension/colloid) forms.

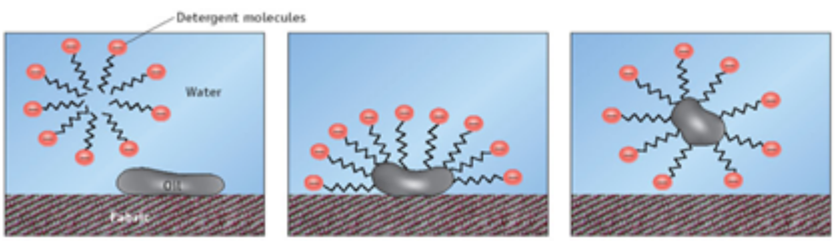
Remember that _____________ dissolves ______________.

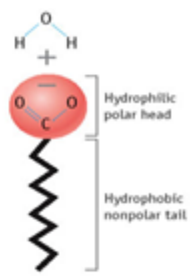


Hydrophilic –

Hydrophobic –

Emulsifying Agent – (like soap) –





Recall: If solute is ionic, ______________________ will form.

In this case, $\Delta H_{soln} = \Delta H_{hydration} - \Delta H_{lattice}$

<u>Practice Problem:</u>

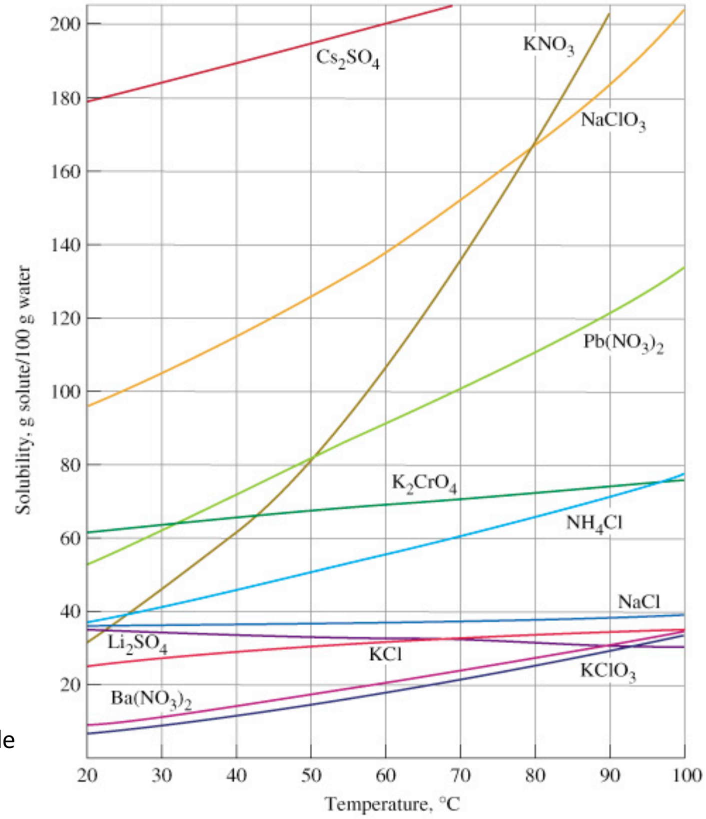
Given $\Delta H^{\circ}_f(s) = -425.9$kJ/mol and $\Delta H^{\circ}_f(aq, 1M) = -469.2$kJ/mol for sodium hydroxide, calculate the enthalpy of a solution of NaOH.

Lesson 3

Most ionic salts become ______________ soluble at higher temperatures.

Saturated Solution–

Supersaturated Solution –



If you heat a solution, you can dissolve more salts. Slowly cooling it will keep those salts in solution. Then, you have a supersaturated solution.

Supersaturated solutions are very fragile. Bumping them or adding 1 tiny piece of solid (________ crystal) can cause the salt to "undissolve" or come out of solution.

Practice Problem:

Using the graph to the right - At 75°C, what is the maximum molarity possible for an aqueous solution of sodium chlorate?

Gases become _____________ soluble at higher temperatures. This is known as _____________ law.

Recall that vapor pressure INCREASES with increasing temperature, so gases become LESS soluble with increasing temperature.

If you want to make a gas MORE soluble, you ________________ the partial pressure of the gas.

Examples: Soda Pop

The Bends

Lake Cameroon

SOLUTIONS

A solution is a **mixture** (usually involving water and another substance). The other substance can be a gas, solid, or another liquid. In a solution, the substances are **evenly**.

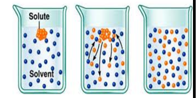
TABLE 14.1 Some Common Solutions

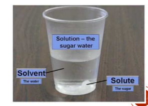
Solution	Components
Gaseous solutions	
Air	N_2, O_2, and several others
Natural gas	CH_4, C_2H_6, and several others
Liquid solutions	
Seawater	H_2O, NaCl, and many others
Vinegar mL of	H_2O, $HC_2H_3O_2$ (acetic acid)
Soda pop	H_2O, CO_2, $C_{12}H_{22}O_{11}$ (sucrose), and several others
Solid solutions	
Yellow brass	Cu, Zn
Palladium–hydrogen	Pd, H_2

VOCABULARY

Homogeneous mixture: **evenly** mixed
Heterogeneous mixture: **unevenly** mixed

Solute – **stuff being dissolved**
Solvent – **substance in greatest quantity**





TYNDALL EFFECT

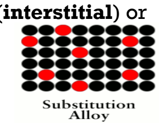
Suspensions/Colloids: **heterogeneous** mixtures;
Suspensions settle out
Colloids don't settle

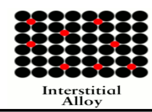
Tyndall Effect – is **light** scattered or not?

ALLOYS

Alloys – mixtures of a metal and a nonmetal (**interstitial**) or 2 metals (**substitutional**).

Alloys – NOT **bonded**;
The mixture has unique properties.





RECALL: MOLARITY

Solutions have different Concentrations! Concentrations are measured in **Molarity** (**M**).

$$6M = \frac{6 \text{ mol}}{1 \text{ L}}$$

PRACTICE PROBLEM:

How many grams of NaCl are needed to make 0.500 L of 0.25M NaCl?

$$\frac{0.500 \cancel{L}}{} \left| \frac{0.25 \cancel{mol}}{1 \cancel{L}} \right| \frac{58.44 \text{ g}}{1 \cancel{mol}}$$

0.25*M* = 0.25 mol / 1L

= 7.3 g NaCl

Things to remember from this problem:
*Watch out for volumes in milliliters!

RECALL: MOLE FRACTIONS

Mole Fractions – a ratio of moles; it's always moles of solute divided by **total** moles

$$X_A = \frac{\text{moles A}}{\text{Total moles}}$$

PRACTICE PROBLEM:

If 10.5g of sodium chloride and 12.6g of potassium chloride are dissolved in 250.g of water, what is the mole fraction of the chloride ion?

$$\frac{10.5\,\cancel{g}}{} \left| \frac{1\text{ mol}}{58.44\,\cancel{g}} \right. = 0.180\text{ mol }Cl^-$$

$$\frac{12.6\,\cancel{g}}{} \left| \frac{1\text{ mol}}{74.55\,\cancel{g}} \right. = 0.169\text{ mol }Cl^-$$

$$\frac{250.\,\cancel{g}}{} \left| \frac{1\text{ mol}}{18.02\,\cancel{g}} \right. = 13.9\text{ mol }H_2O$$

$$\frac{(0.180 + 0.169)}{(0.180 + 0.169 + 13.9)} = 0.0245$$

Things to remember from this problem:
*Moles of ONE substance divided by TOTAL moles.
*YOU CANNOT MAKE A **GRAM** FRACTION!

PRACTICE PROBLEM:

The density of acetonitrile, CH_3CN, is 0.786 g/mL. and the density of methanol, CH_3OH, is 0.791 g/mL. A solution is made by dissolving 20.0mL CH_3OH in 100.0mL of CH_3CN. The vapor pressure of pure CH_3CN is 81.0 mmHg at 22°C.

a. What is the mole fraction of methanol in the solution?

$$\frac{20.0\,\cancel{mL}}{} \left| \frac{0.791\,\cancel{g}}{1\,\cancel{mL}} \right| \frac{1\text{ mol}}{32.05\,\cancel{g}} = 0.494\text{ mol }CH_3OH$$

$$\frac{100.0\,\cancel{mL}}{} \left| \frac{0.786\,\cancel{g}}{1\,\cancel{mL}} \right| \frac{1\text{ mol}}{41.06\,\cancel{g}} = 1.914\text{ mol }CH_3CN$$

$$\frac{0.494}{(0.494 + 1.914)} = 0.205$$

Things to remember from this problem:
*Moles of ONE substance divided by TOTAL moles.

PRACTICE PROBLEM:

The density of acetonitrile, CH_3CN, is 0.786 g/mL. and the density of methanol, CH_3OH, is 0.791 g/mL. A solution is made by dissolving 20.0mL CH_3OH in 100.0mL of CH_3CN. The vapor pressure of pure CH_3CN is 81.0 mmHg at 22°C.

b. What is the new vapor pressure of the CH_3CN with the methanol in solution at 22°C? (Ignore methanol's vapor pressure).

1 - 0.205 = 0.795 This is the mole fraction of CH_3CN.

0.795 x 81.0 = 64.4 mmHg

Things to remember from this problem:
*Apply mole fractions like they are percentages.

PRACTICE PROBLEM:

The density of acetonitrile, CH_3CN, is 0.786 g/mL. and the density of methanol, CH_3OH, is 0.791 g/mL. A solution is made by dissolving 20.0mL CH_3OH in 100.0mL of CH_3CN. The vapor pressure of pure CH_3CN is 81.0 mmHg at 22°C.

c. Assuming that the volumes are additive, what is the molarity of the solution?

*From previous problem: 0.494 mol CH_3OH
*From previous problem: 1.914 mol CH_3CN

$$100.0\text{ mL} + 20.0\text{ mL} \rightarrow \frac{0.494\text{ mol}}{0.120\text{ L}} = 4.12\text{ M}$$

Things to remember from this problem:
*Molarity is moles of SOLUTE divided by TOTAL Liters of solution.

PRACTICE PROBLEM:

A solution was prepared by dissolving 23.7g of $CaCl_2$ in 375g of water. The density of the resulting solution was 1.05g/mL. What is the molarity of the Cl^- ion in this solution? Assume 100% dissociation of the compound. The molar mass of $CaCl_2$ is 110.984g/mol.

$$\frac{23.7\,\cancel{g}}{} \left| \frac{1\,\cancel{mol}}{110.984\,\cancel{g}} \right| \frac{2\text{ mol }Cl^-}{1\,\cancel{mol}\,CaCl_2} = 0.427\text{ mol }Cl^-$$

$$\frac{398.75\,\cancel{g}}{} \left| \frac{1\,\cancel{mL}}{1.05\,\cancel{g}} \right| \frac{1\text{ L}}{1000\,\cancel{mL}} = 0.380\text{ L}$$

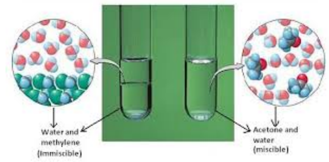
$$\frac{0.427\text{ mol}}{0.380\text{ L}} = 1.12\text{ M}$$

Things to remember from this problem:
*Remember that subscripts are a mole ratio!

LESSON 2

VOCABULARY

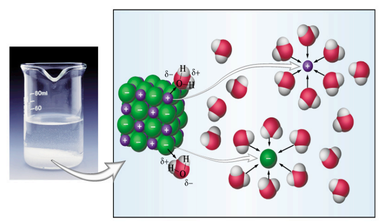
Miscible - Liquids that **mix** with other liquids
Immiscible - Liquids that **don't mix** with other liquids



HOW DO THINGS DISSOLVE?

What is really happening when a solution is made?
1) Solute particles must break apart (overcome intermolecular forces); this is **endothermic** (requires energy)
2) Solvent particles must break apart (overcome intermolecular forces); this is **endothermic** (requires energy)
3) Solute particles & solvent particles mix and form new intermolecular forces; this is **exothermic** (releases energy)
4) The sum of these energy changes is known as enthalpy of solution (**ΔH_{soln}**)

HOW DO THINGS DISSOLVE?



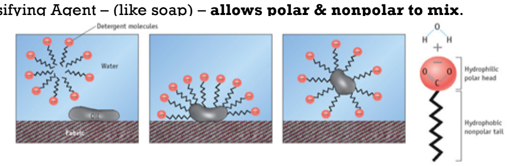
IDEAL & NON-IDEAL SOLUTIONS

Ideal Solution – dissolving causes **no** energy change $\Delta H_{soln} = 0$

Non-ideal Solutions:
$\Delta H_{soln} < 0$; **exo**thermic; intermolecular forces between solute-solvent are **stronger** than the IMF within the solute and the IMF within the solvent.
$\Delta H_{soln} > 0$; **endo**thermic; intermolecular forces between solute-solvent are **a little weaker** than the IMF within the solute and the IMF within the solvent; solution may not form.
$\Delta H_{soln} > >0$; VERY **endo**thermic; intermolecular forces between solute-solvent are MUCH **weaker** than the IMF within the solute and the IMF within the solvent; solution DOES NOT form…heterogeneous mixture (suspension/colloid) forms.

VOCABULARY

Remember that **like dissolves like.**
Hydrophilic – **water-loving (polar)**
Hydrophobic – **water-hating (nonpolar)**
Emulsifying Agent – (like soap) – **allows polar & nonpolar to mix**.



CALCULATING ΔH_{SOLN}

Recall: If solute is ionic, **hydration spheres** will form.
In this case,

$\Delta H_{soln} = \Delta H_{hydration} - \Delta H_{lattice}$

PRACTICE PROBLEM:

Given $\Delta H^{\circ}_{f}(s)$ = -425.9kJ/mol and $\Delta H^{\circ}_{f}(aq, 1M)$ = -469.2kJ/mol for sodium hydroxide, calculate the enthalpy of a solution of NaOH.

$\Delta H_{soln} = \Delta H_{hydration} - \Delta H_{lattice}$

$\Delta H_{soln} = -469.2 - {}^{-}425.9$

= -43.3 kJ

OR

ΔH_{soln} = **Bonds Broken- Bonds Formed**

$\Delta H_{soln} = 425.9 - 469.2$

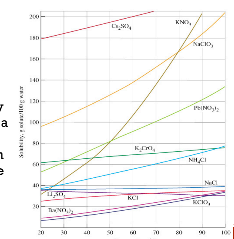
= -43.3 kJ

Things to remember from this problem:
*Identify lattice energy and hydration energy first.
*Hydration – Lattice
*Same as bonds broken – bonds formed!

LESSON 3

DISSOLVING IONIC SALTS

Most ionic salts become **more** soluble at higher temperatures.
Saturated Solution– **maximum solid dissolved**
Supersaturated Solution – **more than maximum solid dissolved**
If you heat a solution, you can dissolve more salts. Slowly cooling it will keep those salts in solution. Then, you have a supersaturated solution.
Supersaturated solutions are very fragile. Bumping them or adding 1 tiny piece of solid (**seed** crystal) can cause the salt to "undissolve" or come out of solution.

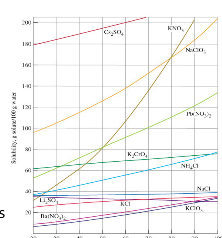


PRACTICE PROBLEM:

Using the graph to the right - At 75°C, what is the maximum molarity possible for an aqueous solution of sodium chlorate?

$$\frac{160\ \cancel{g}}{} \left| \frac{1\ mol}{106.44\ \cancel{g}} \right. = 1.50\ mol\ NaClO_3$$

$$\frac{\mathbf{1.50\ mol}}{\mathbf{0.100\ L}} = 15\ M$$



Things to remember from this problem:
*Notice there are a very few ionic salts that become less soluble with increasing temperature.

DISSOLVING GASES

Gases become **less** soluble at higher temperatures. This is known as **Henry's** law.
Recall that vapor pressure INCREASES with increasing temperature, so gases become LESS soluble with increasing temperature.

If you want to make a gas MORE soluble, you **increase** the pressure on the gas.

Examples: Soda Pop
The Bends
Lake Cameroon

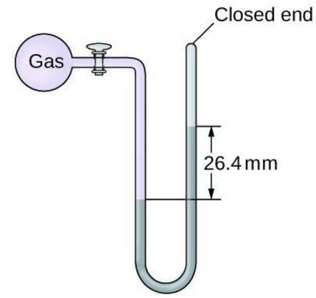
Gas Laws Guided Notes

Pressure of a Gas

- Pressure is the ____________ that a gas puts on its container.
- (Based on the frequency of gas molecule ____________ with the walls of its container).
- Measured in ____________.

Units for Pressure

- ____________ **mmHg =** ____________ **atm**
- ____________ **mmHg =** ____________ **torr**
- ____________ **mmHg =** ____________ **kPa**

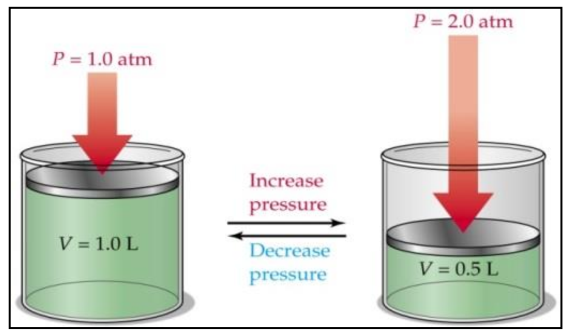


Practice Problem:

- Convert 799 mmHg into atm.

Boyle's Law

- If **pressure** ________________, **volume** ____________.



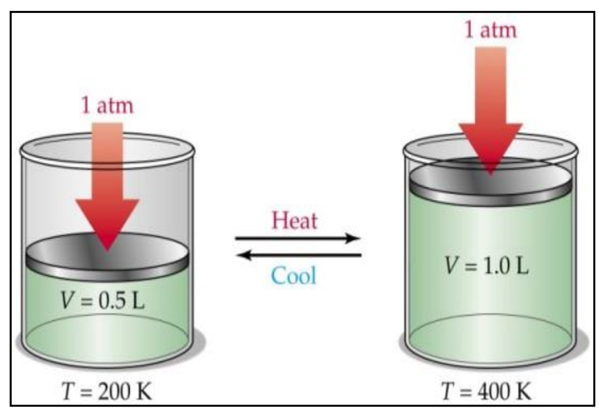
$$P_1V_1 = P_2V_2$$

Practice Problem:

- One car tire has a volume of 10.0 Liters of air at 1520 mmHg. When it's attached to the car, the weight of the car compresses the tire to a volume of 9.5 Liters. What's the new pressure of the tire attached to the car?

Charles' Law

- If **temperature** ____________, **volume** ____________.
- Temperature MUST BE in ____________!
- **Kelvin = Celsius +** ____________



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

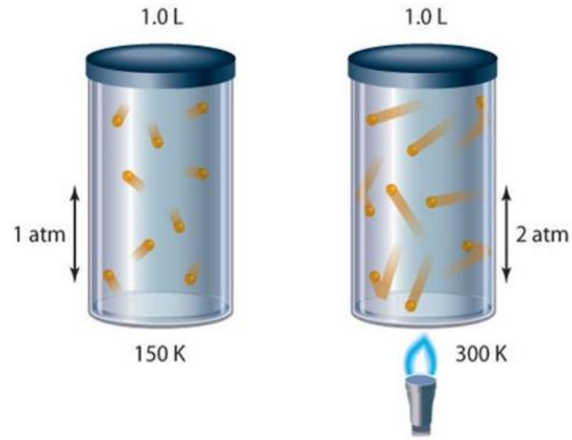
Practice Problem:

- A car tire has a volume of 10.0 Liters of air at 22°C. In the winter, when the temperature is only 1.0°C, what is the volume of the car tire?

Gay-Lussac's Law

- If **temperature ____________, pressure ____________**.
- Temperature MUST BE in ____________**!**
- **Kelvin = Celsius + ____________**

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



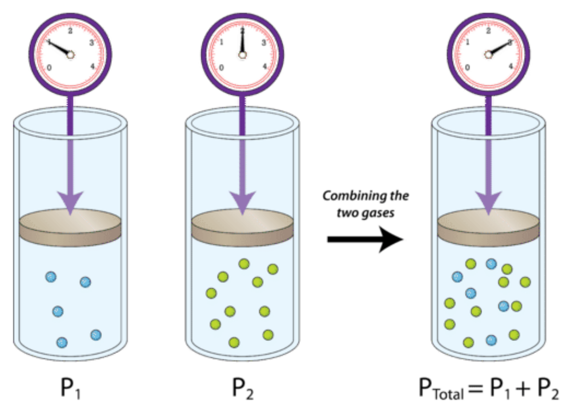
Practice Problem:

- A car tire has a pressure of 1520. mmHg of air at 22.00°C. After driving for a while, the tire has heated up to 120.0°C. What is the pressure of the hot car tire?

Dalton's Law of Partial Pressures

- Combining gases will ____.
- Individual pressures are called ____________________.
- When combined, they form the ____________________.

$$P_{total} = P_1 + P_2 \ldots.$$



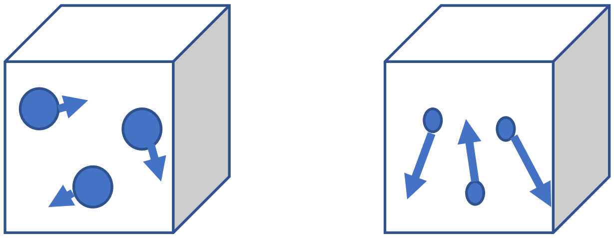
Volume and temperature are constant

Practice Problem:

- During surgery, oxygen, at 121.6 mmHg, is combined with anesthesia at 638.4 mmHg. What is the total pressure of gas given to this patient?

Mass of a Gas Affects Speed

- _____________ _____________ _____________ _____________ (lower speed based on molar mass).



Practice Problem:

- Oxygen and Neon are placed in the same container at 273K. Which gas is moving faster?

Gases are Collected over Water

- Water has a vapor above it.
- This water vapor has a pressure.
- So the pressure in the collection vessel is BOTH the pressure of the gas and the pressure of the water vapor.
- _____________ _____________ _____________ _____________ _____________ from the total pressure!

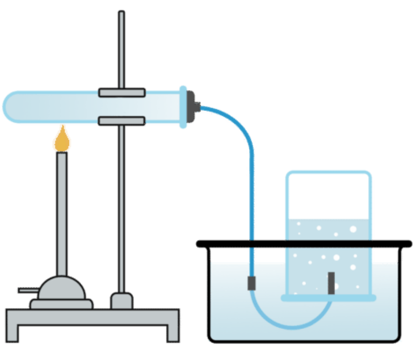
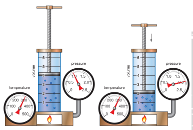


TABLE 10.2 VAPOR PRESSURE OF WATER

TEMPERATURE (°C)	PRESSURE (mm Hg)	TEMPERATURE (°C)	PRESSURE (mm Hg)
5	6.5	55	118.0
10	9.2	60	149.4
15	12.8	65	187.5
20	17.5	70	233.7
25	23.8	75	289.1
30	31.8	80	355.1
35	41.2	85	433.6
40	55.3	90	525.8
45	71.9	95	633.9
50	92.5	100	760.0

Practice Problem:

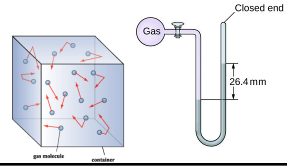
- CO_2 gas is collected by water displacement at 40°C. If the pressure on the collection vessel is 395.0 mmHg, what is the pressure of the CO_2? (The vapor pressure of water at 40°C is 55.3 mmHg).



Gas Laws

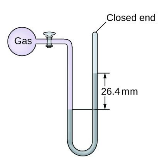
PRESSURE OF A GAS

- Pressure is the **force** that a gas puts on its container.
- (Based on the frequency of gas molecule **collisions** with the walls of its container).
- Measured in **mmHg.**



UNITS FOR PRESSURE

- **760 mmHg = 1 atm**
- **760 mmHg = 760 torr**
- **760 mmHg = 101.325 kPa**



PRACTICE PROBLEM:

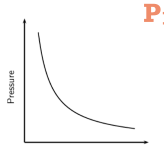
- Convert 799 mmHg into atm.

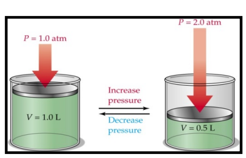
$$\frac{799\ \text{mmHg}}{} \left| \frac{1\ \text{atm}}{760\ \text{mmHg}} \right. = 1.05\ \text{atm}$$

BOYLE'S LAW

- If **pressure increases, volume decreases.**

$$P_1V_1 = P_2V_2$$





PRACTICE PROBLEM:

- One car tire has a volume of 10.0 Liters of air at 1520 mmHg. When it's attached to the car, the weight of the car compresses the tire to a volume of 9.5 Liters. What's the new pressure of th tire attached to the car?

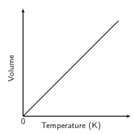
$$P_1V_1=P_2V_2$$

$$(1520)(10.0) = P(9.5)$$

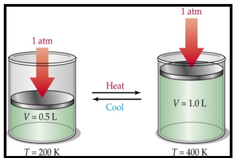
$$P = 1600\ \text{mmHg}$$

CHARLES' LAW

- If **temperature increases, volume increases.**
- Temperature MUST BE in **Kelvin!**
- **Kelvin = Celsius + 273**



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



PRACTICE PROBLEM:

- A car tire has a volume of 10.0 Liters of air at 22°C. In the winter, when the temperature is only 1.0°C, what is the volume of the car tire?

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

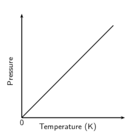
(10.0)÷ (295) = V÷ (274)

V = 9.288 → 9 .3 L

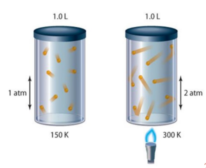
TEMP MUST B E IN

GAY-LUSSAC'S LAW

- If **temperature increases, pressure increases.**
- Temperature MUST BE in **Kelvin!**
- **Kelvin = Celsius + 273**



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



PRACTICE PROBLEM:

- A car tire has a pressure of 1520. mmHg of air at 22.00°C. After driving for a while, the tire has heated up to 120.0°C. What is the pressure of the hot car tire?

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

(1520)÷ (295) = P ÷ (393)

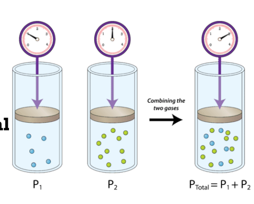
P = 2024.949 → 2025 mmHg

TEMP MUST B E IN

DALTON'S LAW OF PARTIAL PRESSURES

- Combining gases will **combine their pressures.**
- Individual pressures are called **partial pressures.**
- When combined, they form the **total pressure.**

$P_{total} = P_1 + P_2$....



PRACTICE PROBLEM:

- During surgery, oxygen, at 121.6 mmHg, is combined with anesthesia at 638.4 mmHg. What is the total pressure of gas given to this patient?

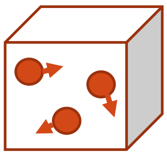
$P_{total} = P_1 + P_2$....

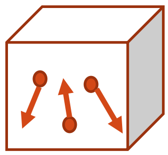
121.6 + 638 .4 = 760 mmHg

MASS OF A GAS AFFECTS SPEED

- **Heavier gases move slower** (lower speed based on molar mass).

Heavier = Slower **Lighter = Faster**





PRACTICE PROBLEM:

- Oxygen and Neon are placed in the same container at 273K. Which gas is moving faster?

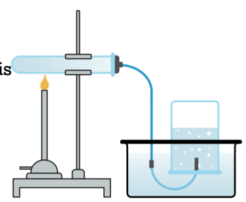
Oxygen = O_2 = 32.00 g/mol
Neon = Ne = 20 g/mol

Neon is lighter.

Neon is moving faster.

GASES ARE COLLECTED OVER WATER

- Water has a vapor above it.
- This water vapor has a pressure.
- So the pressure in the collection vessel is BOTH the pressure of the gas and the pressure of the water vapor.
- **Subtract the water vapor pressure** from the total pressure!



PRACTICE PROBLEM:

- CO_2 gas is collected by water displacement at 40°C. If the pressure on the collection vessel is 395.0 mmHg, what is the pressure of the CO_2? (The vapor pressure of water at 40°C is 55.3 mmHg).

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35	41.2	85	433.6
40	55.3	90	525.8
45	71.9	95	633.9
50	92.5	100	760.0

395.0 – 55.3 = **339 .7 mmHg**