OXIDATION / REDUCTION CHEMISTRY

- Exchange reactions involve ions pairing up, but the ions themseves are not formed in exchange reactions. Exchanges start with pre-existing ions.

... but the ions have to be produced somehow - through a chemistry that involves the transfer of electrons.

- oxidation / reduction chemistry ("redox" chemistry) involves transfer of electrons and can make ions.

$$2Al(s) + 3Br_2(l) \rightarrow 2AlBr_3(s)$$

Elemental, metallic aluminum. Uncharged!

Al^0

Aluminum cation

Al^{3+}

These are called "half-reactions"

$$Al \rightarrow Al^{3+} + 3\underbrace{e^-}_{\text{electron}}$$

oxidation: loss of electrons

$$Br_2 + 2e^- \rightarrow 2Br^-$$

reduction: gain of electrons

- oxidation and reduction always occur together. In other words, we can't just make free electrons using oxidation without giving them somewhere to go.

- Many of the types of reactions that we learned about in previous courses are redox reactions!
  - COMBINATIONS (often but not always redox)
  - DECOMPOSITIONS (often redox)
  - SINGLE REPLACEMENT (always redox)

$$Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

$$Cu \rightarrow Cu^{2+} + 2e^- \quad \text{oxidation}$$

$$2Ag^+ + 2e^- \rightarrow 2Ag(s) \quad \text{reduction}$$

net ionic → $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

- COMBUSTION

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

$$2Mg(s) \rightarrow 2Mg^{2+} + 4e^- \quad \text{oxidation}$$

$$O_2(g) + 4e^- \longrightarrow 2O^{2-} \quad \text{reduction}$$

A review of the reaction types we just mentioned:

① COMBINATION REACTIONS

- Reactions that involve two or more simple substances COMBINING to form a SINGLE product

- Often involve large energy changes. Sometimes violent!

- Form: $\boxed{A + B + \ldots \longrightarrow C}$

Example:

$$2\,Al(s) + 3\,Br_2(l) \longrightarrow 2\,AlBr_3(s)$$

② DECOMPOSITION REACTIONS

- Reactions where a SINGLE REACTANT breaks apart into several products

- Form: $\boxed{A \longrightarrow B + C + \ldots}$

Example:

$$2\,H_2O_2(l) \longrightarrow 2\,H_2O(l) + O_2(g)$$

* This reaction is NOT a combustion reaction, even though O_2 is involved!

* Combustion reactions CONSUME O_2, while this reaction PRODUCES O_2

③ COMBUSTION REACTIONS

- Reactions of substances with MOLECULAR OXYGEN (O_2) to form OXIDES.

- Combustion forms an OXIDE of EACH ELEMENT in the burned substance!

- Form: $\boxed{AB + O_2 \longrightarrow AO + BO}$

Oxide: a compound containing OXYGEN and one other element!

* Combustion of hydrocarbons makes carbon dioxide and water, if enough oxygen is present. In low-oxygen environments, carbon monoxide is made instead!

Examples:

Oxides!

*

$$C_3H_8(g) + 5O_2(g) \longrightarrow 4H_2O(g) + 3CO_2(g)$$

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

This reaction can also be called a combination! Two reactants form a single product.

④ SINGLE REPLACEMENT REACTIONS

- Reactions where one element REPLACES another element in a compound.

- Can be predicted via an ACTIVITY SERIES (more on that later!)

- Form: $\boxed{A + BC \longrightarrow AC + B}$ "A" and "B" are elements., often metals.

- Easy to spot, since there is an element "by itself" on each side of the equation.

Examples:

$$\overset{0}{Cu}(s) + 2\overset{+1}{Ag}NO_3(aq) \longrightarrow \overset{+2}{Cu}(NO_3)_2(aq) + 2\overset{0}{Ag}(s)$$

$$\overset{0}{Zn}(s) + \overset{+1}{H}_2SO_4(aq) \longrightarrow \overset{+2}{Zn}SO_4(aq) + \overset{0}{H}_2(g)$$

REDOX LANGUAGE

"oxidizer"

- "Oxidation" is loss of electrons, but an OXIDIZING AGENT is something that causes ANOTHER substance to lose electrons. An oxidizing agent is itself reduced during a redox reaction.

- "Reduction" is gain of electrons, but a REDUCING AGENT is something that causes ANOTHER substace to gain electrons. Reducing agents are themselves oxidized during a redox reaction.

$$2\,Al(s) + 3\,Br_2(l) \longrightarrow 2\,AlBr_3(s)$$

Aluminum is OXIDIZED during this process. We say that metallic aluminum is a REDUCING AGENT!

Bromine is REDUCED during this process. We say that bromine is an OXIDIZING AGENT!

* Strong oxidizers (oxidizing agents) can cause spontaneous fires if placed into contact with combustibles (safety issue!).

GASES

- Gases differ from the other two phases of matter in many ways:

  - They have very low viscosity (resistance to flow), so they flow from one place to another very easily.

  - They will take the volume of their container. In other words, gas volumes are variable.

  - They are the least dense of all three phases.

  - Most gases are transparent, and many are invisible.

  thermal expansion

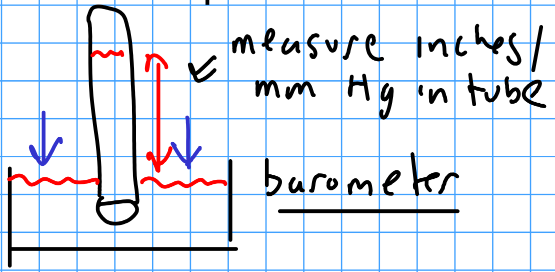
  - Gases show a much larger change of volume on heating or cooling than the other phases.

  - Gases react to changes in temperature and pressure in a very similar way. This reaction often does not depend on what the gas is actually made of.

KINETIC THEORY

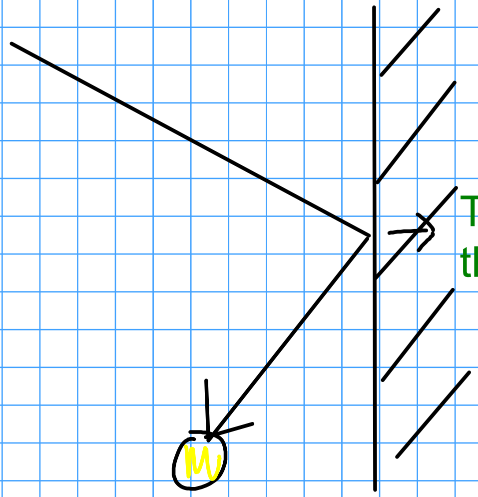
- is a way to explain the behavior of gases.
- views the properties of gases as arising from them being molecules in motion.

- Pressure: force per unit area. Units: Pascal, bar, mm Hg, in Hg, atm, etc.



760 mm Hg = 1 atm

- According to kinetic theory, pressure is caused by collisions of gas molecules with each other and the walls of the container the gas is in.



This collision exerts a force on the wall!

- Temperature:
  - a measure of the average kinetic energy of the molecules of the gas

$$E_K = \frac{1}{2} m v^2$$

velocity

mass

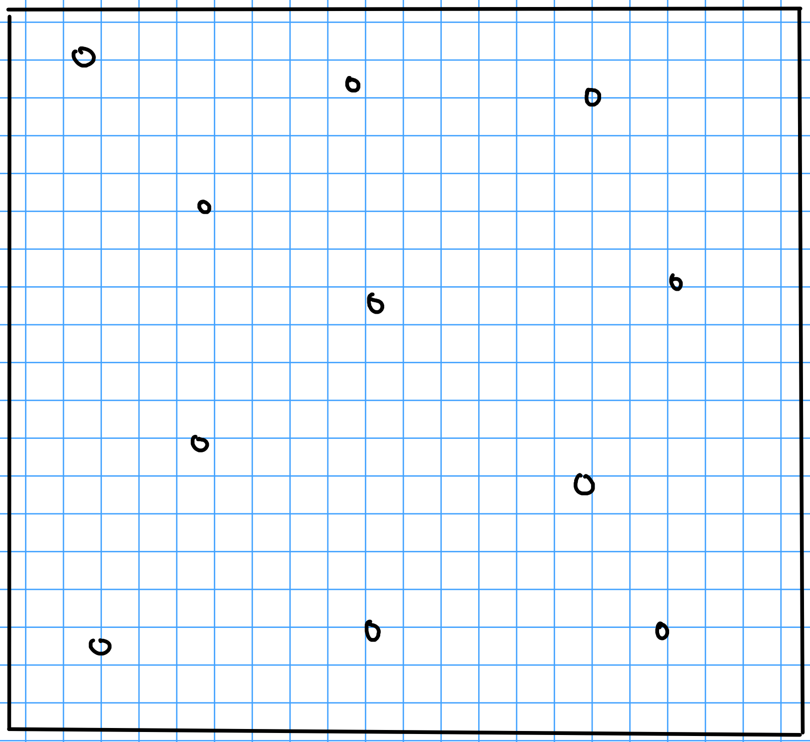
- The faster the gas molecules move, the higher the temperature!
- The temperature scales used when working with gases are ABSOLUTE scales.
  - ABSOLUTE: scales which have no values less than zero.
    - KELVIN: metric absolute temperature scale.

$$\boxed{K = 273.15 + {}^\circ C}$$

Quick comparison of temperature scales!

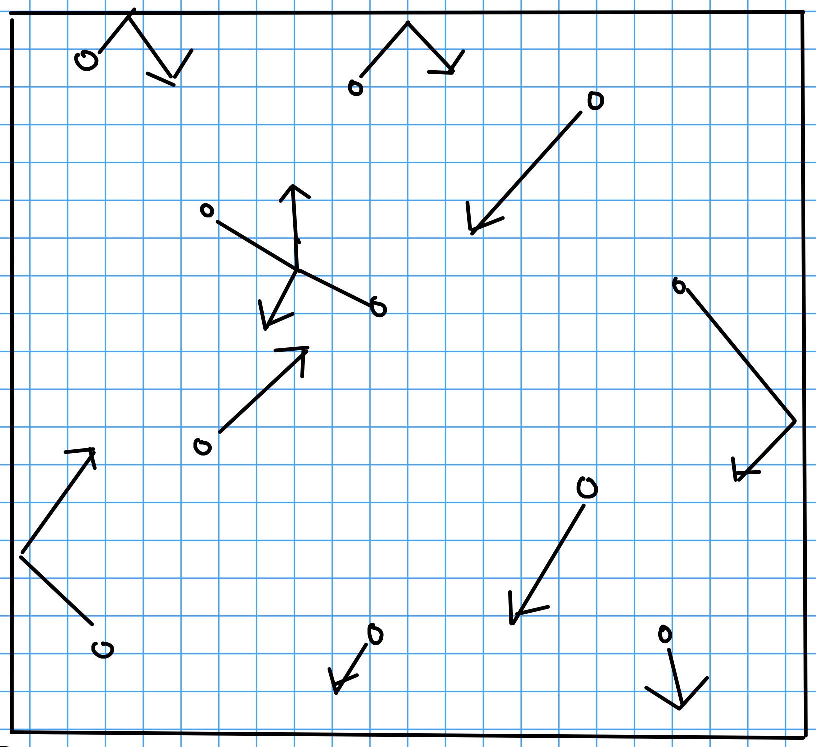
212	100	373	Water boils
77	25	298	Room temperature
32	0	273	Water freezes
-460	-273	0	Absolute zero!
°F	°C	K	

THE KINETIC PICTURE OF GASES



① Gas molecules are small compared to the space between the gas molecules!

LOW DENSITY!



② Gas molecules are constantly in motion. They move in straight lines in random directions and with various speeds.

③ Attractive and repulsive forces between gas molecules are so small that they can be neglected except in a collision.

- Each gas molecule behaves independently of the others.

④ Collisions between gas molecules and each other or the walls are ELASTIC.

⑤ The average kinetic energy of gas molecules is proportional to the absolute temperature.

How does this picture explain the properties of gases?

- Gases expanding to fill their container? Agrees with kinetic picture, since gas molecules are independent
- Thermal expansion of gas at constant pressure? Agrees, because the container has to EXPAND to keep the pressure (from collisions) constant when the gas molecules move faster.
- Pressure increases with temperature at constant volume: Agrees, because the number and force of collisions increases with molecular speed.

GAS LAWS

- were derived by experiment long before kinetic theory, but agree with the kinetic picture!

Boyle's Law:

$$PV = \text{constant}$$

True at constant temperature

$$P_1V_1 = \text{constant} \qquad P_2V_2 = \text{constant}$$

$$\boxed{P_1V_1 = P_2V_2}$$

True at constant temperature

Charles's Law:

$$\frac{V}{T} = \text{constant}$$

True at constant pressure, and
using ABSOLUTE temperature

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

True at constant pressure, and
using ABSOLUTE temperature

Combined gas law:

$$\frac{PV}{T} = \text{constant}$$

Must use ABSOLUTE temperature units!

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

Must use ABSOLUTE temperature units!

amount (moles) of gas must be constant!

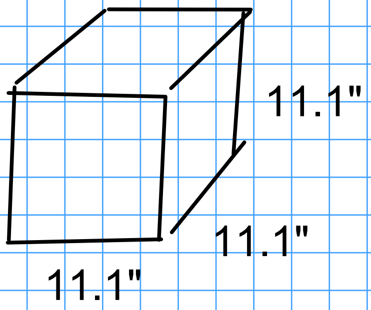
Avogadro's law:

- a mole of any gas at the same conditions has the same volume.

1 mol gas molecules @ 0°C and 1 atm "STP"

Standard
Temperature
and
Pressure

volume = 22.4 L



= 22.4 L

Ideal gas law:

$$\frac{PV}{T} = \underline{\text{constant}}$$

... but this constant actually depends on the amount of gas!

$$\longrightarrow \; = n \times \text{"}R\text{"}$$

The ideal gas constant.

$$0.08206 \; \frac{L \cdot atm}{mol \cdot K}$$

... combining this together ...

$$\frac{PV}{T} = nR$$

$$\downarrow$$

$$\boxed{PV = nRT}$$

P = pressure atm

V = volume L

T = ABSOLUTE temperature K

R = ideal gas constant

n = number of moles of gas molecules

CHEMICAL CALCULATIONS WITH THE GAS LAWS

$$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2H_2O(l) + 2CO_2(g) + Na_2SO_4(aq)$$

Given 25.0 g of sodium bicarbonate and sufficient sulfuric acid, what volume of carbon dioxide gas would be produced at 25.0 C and 0.950 atm pressure?

$FW_{NaHCO_3} = 84.007\ g/mol$

① g bicarb → mol bicarb
② mol bicarb → mol CO_2
③ mol CO_2 → volume CO_2

$$25.0\,g\ NaHCO_3 \times \frac{mol}{84.007\,g} \times \frac{2\ mol\ CO_2}{2\ mol\ NaHCO_3} = 0.2975942\ mol\ CO_2$$

$$PV = nRT$$

$P = 0.950\ atm$

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

$T = 25.0°C = 298.2\ K$

$V = ?$

$$V = \frac{nRT}{P} = \frac{(0.2975942\ mol)\left(0.08206\ \frac{L\cdot atm}{mol\ K}\right)(298.2\ K)}{(0.950\ atm)}$$

$$= \boxed{7.67\ L}$$

What volume would the gas in the last example problem have at STP?

STP: 0°C, 1 atm

7.67 L CO_2 at 0.950 atm and 298.2 K

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P_1 = 0.950 atm

V_1 = 7.67 L

T_1 = 298.2 K

P_2 = 1.00 atm

V_2 = ?

T_2 = 273.2 K

$$\frac{P_1V_1T_2}{T_1P_2} = V_2 = \frac{(0.950\ \text{atm})(7.67\ \text{L})(273.2\ \text{K})}{(298.2\ \text{K})(1.00\ \text{atm})}$$

$$= \boxed{6.68\ \text{L}}$$