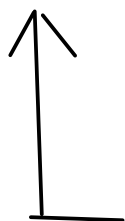
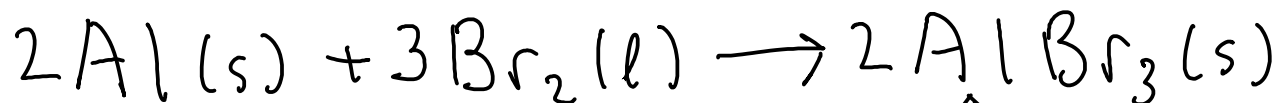


125 OXIDATION / REDUCTION CHEMISTRY

- Exchange reactions involve ions pairing up, but the ions themselves 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.

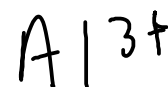
- oxidation / reduction chemistry ("redox" chemistry) involves transfer of electrons and can make ions.



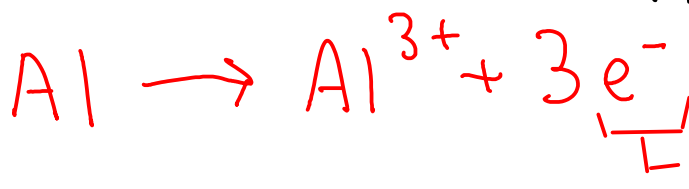
Elemental,
metallic
aluminum.
Uncharged!



Aluminum
cation



These are called
"half-reactions"



electron

oxidation: loss
of electrons



reduction: gain of
electrons

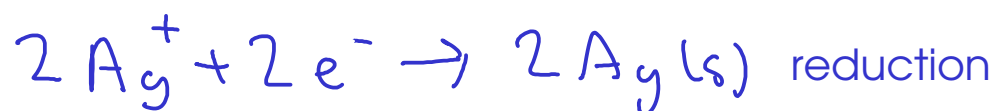
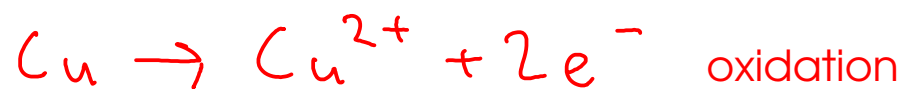
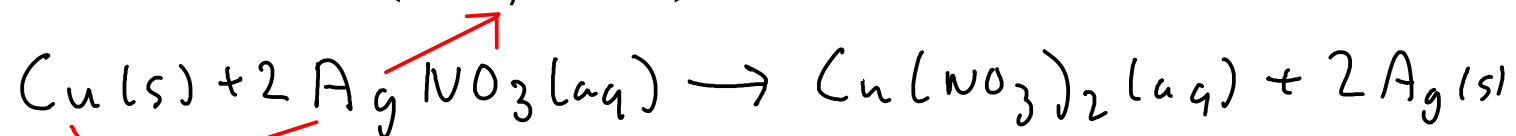
- oxidation and reduction always occur together. In other words, we can't just make free electrons using oxidation without giving them somewhere to go.

- Many of the types of reactions that we learned about in previous courses are redox reactions!

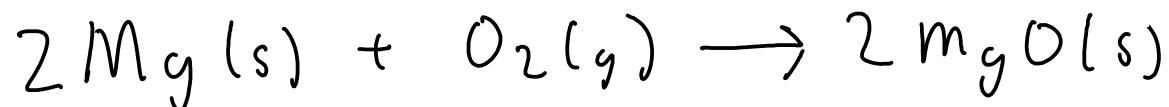
- COMBINATIONS (often but not always redox)

- DECOMPOSITIONS (often redox)

- SINGLE REPLACEMENT (always redox)



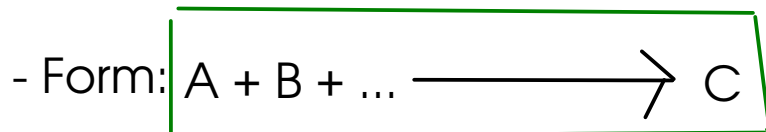
- COMBUSTION



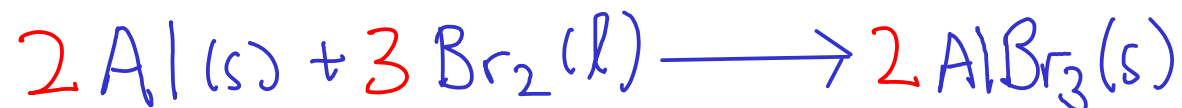
A review of the reaction types we just mentioned:

1) COMBINATION REACTIONS

- Reactions that involve two or more simple substances COMBINING to form a SINGLE product
- Often involve large energy changes. Sometimes violent!

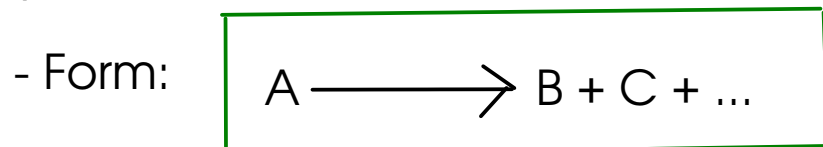


Example:

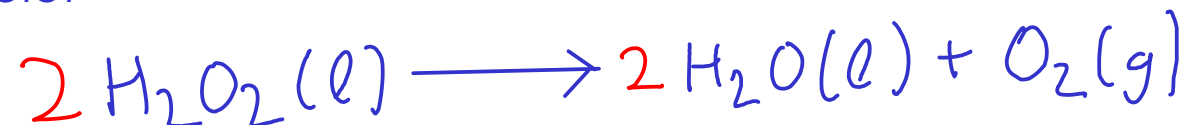


② DECOMPOSITION REACTIONS

- Reactions where a SINGLE REACTANT breaks apart into several products



Example:

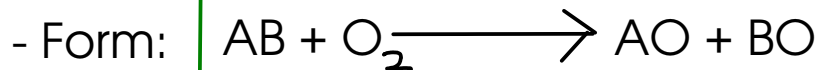


* This reaction is NOT a combustion reaction, even though O_2 is involved!

* Combustion reactions CONSUME O_2 , while this reaction PRODUCES O_2

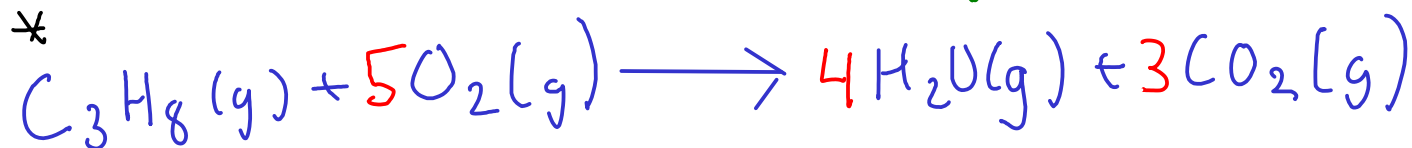
3 COMBUSTION REACTIONS

- Reactions of substances with MOLECULAR OXYGEN (O_2) to form OXIDES.
- Combustion forms an OXIDE of EACH ELEMENT in the burned substance!



Oxide: a compound containing OXYGEN and one other element!

Examples:



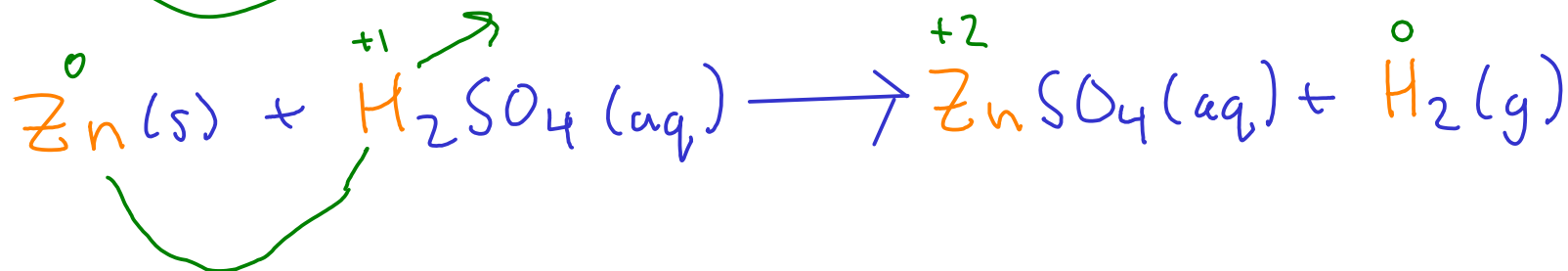
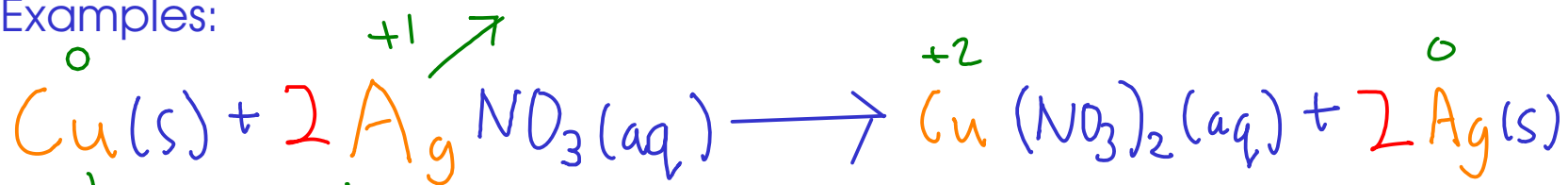
This reaction can also be called a combination!
Two reactants form a single product.

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

4 SINGLE REPLACEMENT REACTIONS

- Reactions where one element REPLACES another element in a compound.
- Can be predicted via an ACTIVITY SERIES (p151, 9th edition) (p153, 10th ed)
- Form: $A + BC \longrightarrow AC + B$ "A" and "B" are elements., often metals.
- Easy to spot, since there is an element "by itself" on each side of the equation.

Examples:

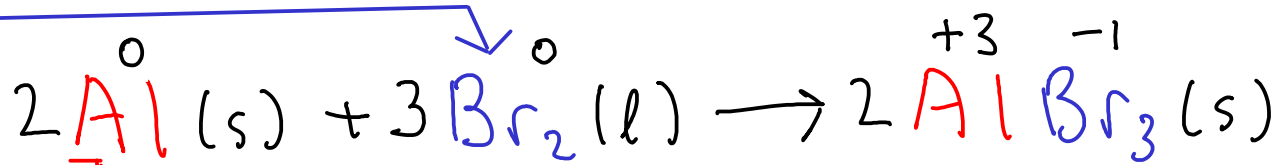


REDOX LANGUAGE

"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 substance to gain electrons. Reducing agents are themselves oxidized during a redox reaction.



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!).

* Reactive metals tend to be REDUCING AGENTS, while oxygen-rich ions like NITRATES tend to be OXIDIZING AGENTS. HALOGENS (Group VIIA) also tend to be OXIDIZING AGENTS

END OF CHAPTER 4 MATERIAL

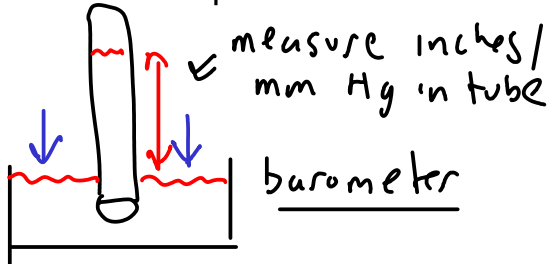
GASES

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

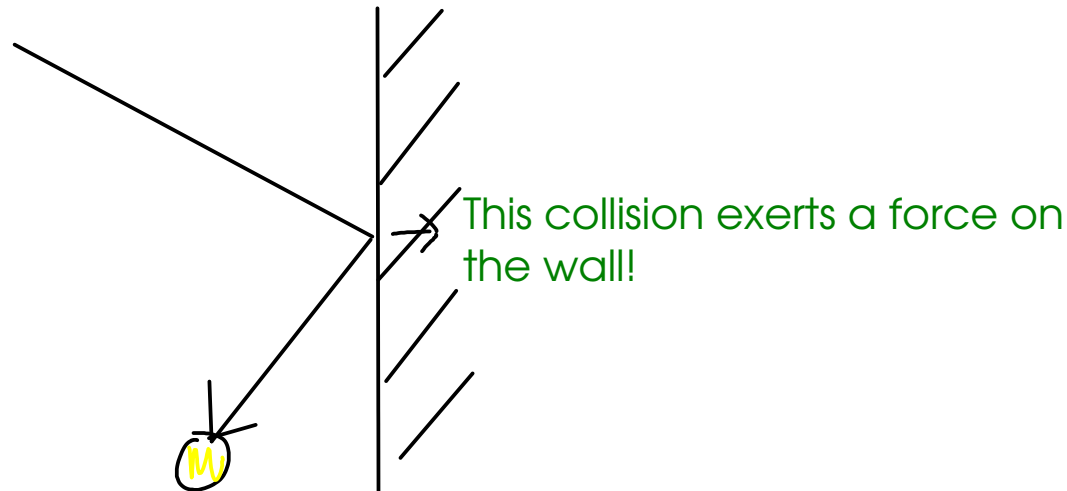
- is a way to explain the behavior of gases.
 - views the properties of gases as arising from them being molecules in motion.
-

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



$$760 \text{ mm Hg} = 1 \text{ atm}$$

- 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.



135- Temperature:

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

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

velocity
mass

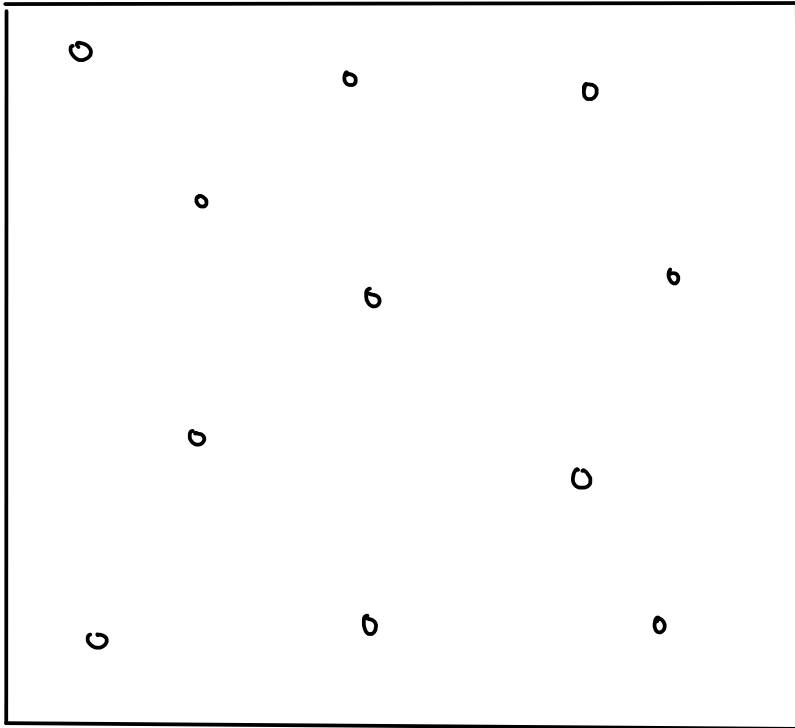
- The faster the gas molecules move, the higher the temperature!
- The temperature scales used when working with gases are ABSOLUTE scales.
 - ABSOLUTE: scales which have no values less than zero.
 - KELVIN: metric absolute temperature scale.

Quick comparison of temperature scales!

$$K = 273.15 + ^\circ C$$

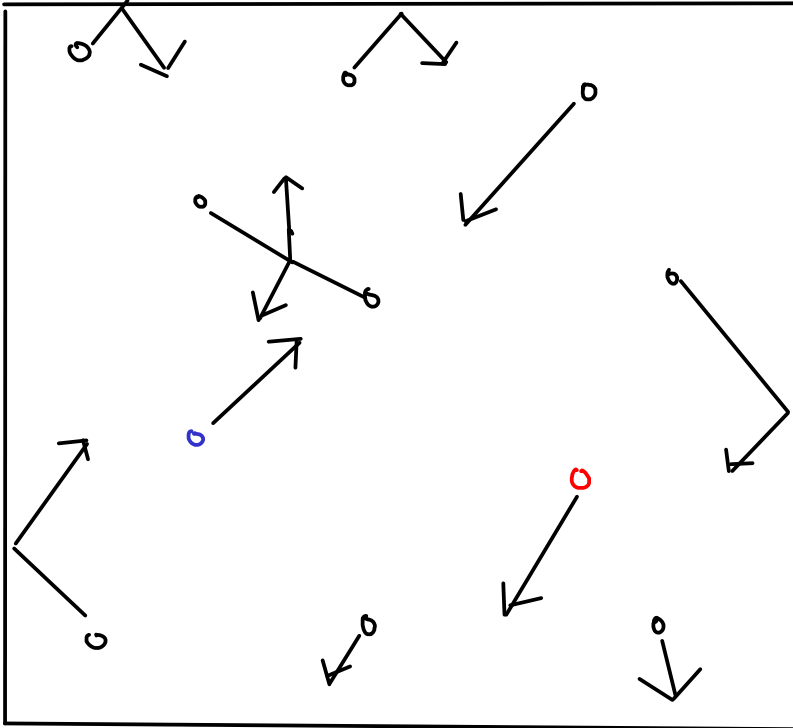
212	100	373	Water boils
77	25	298	Room temperature
32	0	273	Water freezes
-460	-273	0	Absolute zero!
$^{\circ}F$	$^{\circ}C$	K	

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:

$$PV = \text{constant} \quad \left. \vphantom{PV = \text{constant}} \right] \text{ True at constant temperature}$$

$$P_1 V_1 = \text{constant}$$

$$P_2 V_2 = \text{constant}$$

$$\left. \vphantom{P_1 V_1 = \text{constant}} \right] \rightarrow \boxed{P_1 V_1 = P_2 V_2} \quad \text{True at constant temperature}$$

Charles's Law:

$$\frac{V}{T} = \text{constant} \quad \left. \vphantom{\frac{V}{T} = \text{constant}} \right] \text{ True at constant pressure, and using ABSOLUTE temperature}$$

$$\left. \vphantom{\frac{V}{T} = \text{constant}} \right] \rightarrow \boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}} \quad \text{True at constant pressure, and using ABSOLUTE temperature}$$

Combined gas law:

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

Must use ABSOLUTE temperature units!

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

Must use ABSOLUTE temperature units!

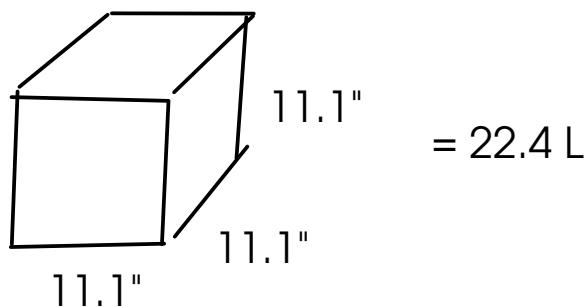
↑ amount (moles) of gas must be constant!

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
 volume = 22.4 L

"STP"
 Standard
 Temperature
 and
 Pressure



Ideal gas law:

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

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

$$= n \times "R"$$

The ideal gas constant.

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

... combining these together ...

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



$$PV = nRT$$

P = pressure atm

V = volume L

T = ABSOLUTE temperature K

R = ideal gas constant

n = number of moles of gas molecules