EOC

4A Properties & Changes of Matter

Physical properties: a quality of matter that can be observed or measured without changing it. Examples of physical properties include: color, molecular weight, change of state and volume.

Physical Change: substances may change shapes or forms while undergoing a physical change, but no chemical reactions occur and no new compounds are produced.

Examples of Physical Changes: crumpling a sheet of paper, melting an ice cube, , dissolving salt in water, casting silver in a mold, breaking a bottle

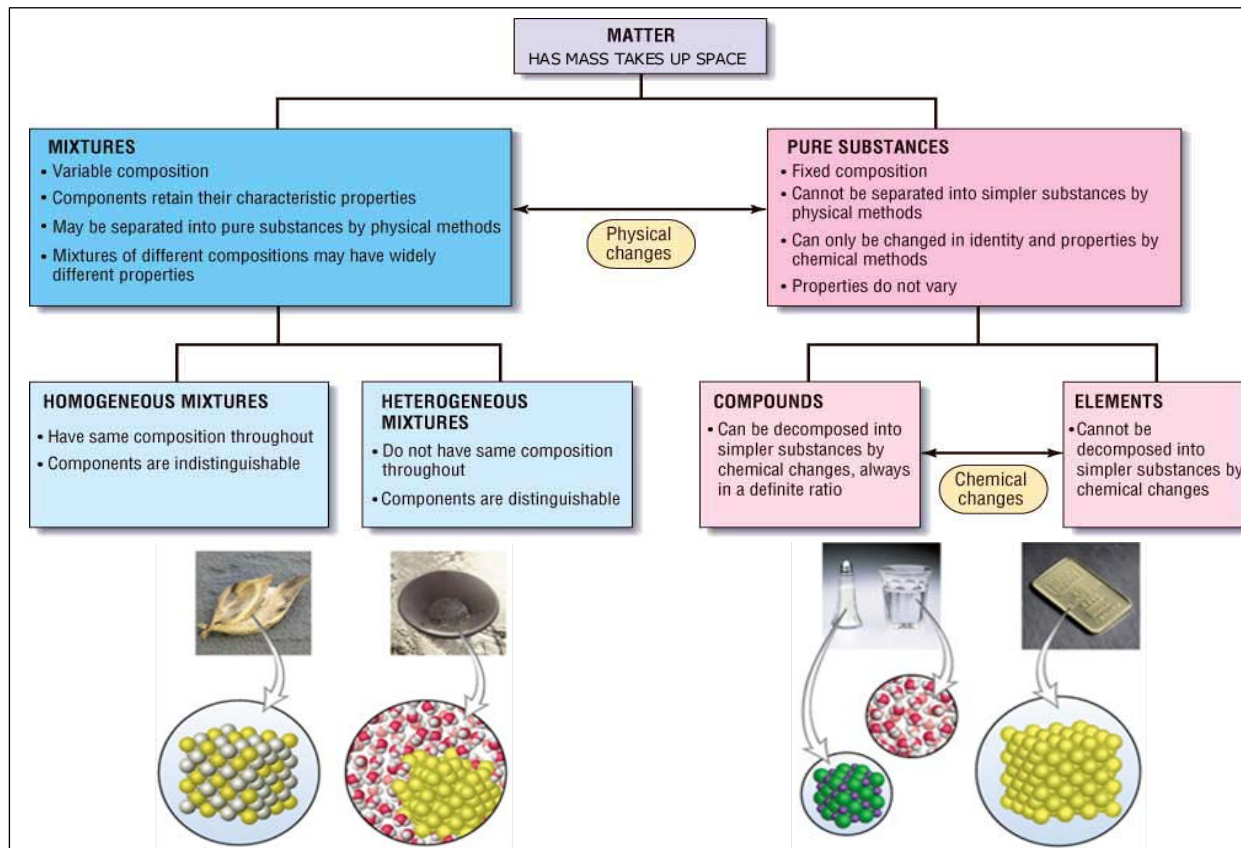**Liquid Water** →(Heat)→ **Water Vapor**

Chemical

Properties: may only be observed by changing the chemical identity of a substance. This property measures the potential for undergoing a chemical change. Examples of chemical properties include reactivity, flammability and oxidation states

Chemical Change: produces a new substance (rearranges the atoms and changes the composition of the starting materials

*Examples of Chemical Changes:* burning wood, mixing acid and base, digesting food, car rusting

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4B   *Extensive Properties*: Depend on the amount of the substance NOT on the identity of the substance - Examples: Mass, Volume Length

Intensive Properties: Depend on the identity but NOT dependent on the amount but - Examples: density, boiling point, solubility, specific heat

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4C States

Gas	Liquid	Solid
• low density • easy to expand/compress / molecules become closer when compresses • fills container/takes volume & shape of container • weak attraction between molecules	• high density • hard to expand/compress • takes shape of container not volume • more attraction between molecules of a gas and less than molecules of a solid • molecules flow yet has a fixed volume	• high density • hard to expand/compress • rigid shape/ does not take shape or volume of container • strong attractions between molecules • molecules in a fixed position

MATTER
HAS MASS TAKES UP SPACE

MIXTURES
- Variable composition
- Components retain their characteristic properties
- May be separated into pure substances by physical methods
- Mixtures of different compositions may have widely different properties

Physical changes

PURE SUBSTANCES
- Fixed composition
- Cannot be separated into simpler substances by physical methods
- Can only be changed in identity and properties by chemical methods
- Properties do not vary

HOMOGENEOUS MIXTURES
- Have same composition throughout
- Components are indistinguishable

HETEROGENEOUS MIXTURES
- Do not have same composition throughout
- Components are distinguishable

COMPOUNDS
- Can be decomposed into simpler substances by chemical changes, always in a definite ratio

Chemical changes

ELEMENTS
- Cannot be decomposed into simpler substances by chemical changes

5A-History of Periodic Table			
Contributor	Date	Contribution	Comment
Lothar Meyer	1869	Compiled a Periodic Table of 56 elements based on the **periodicity of properties such as molar volume when arranged in order of atomic weight.**	Meyer & Mendeleev produced their Periodic Tables simultaneously.
Dmitri Mendeleev	1869	He wrote out the properties of each element on a different card and spent a great deal of time arranging and rearranging them. He was looking for patterns or trends in the data on the cards. Mendeleev first arranged all the cards from lowest to highest atomic mass. He kept elements with similar chemical and physical properties grouped in families. His genius is shown in his ability to recognize the potential for missing data and to use existing data to predict the properties of these unknown elements. **Mendeleev wrote the following Periodic Law: the properties of the elements are a periodic function of their atomic weights.**	Mendeleev's Periodic Table was important because it enabled the properties of elements to be predicted by means of the 'periodic law': properties of the elements vary **periodically with their atomic weights.**
Henry Moseley	1913	In 1914, **Henry Moseley determined the atomic number of each of the known elements**. He realized that, if the elements were arranged in order of increasing atomic number rather than atomic weight, they gave a better fit within the 'periodic table'. He puts the elements tellerium and iodine, argon and potassium, cobalt and nickel in the right order.	**He modified the 'Periodic Law' to read that the properties of the elements vary periodically with their atomic numbers.**

5B-**Families** are groups of elements that react in similar ways.
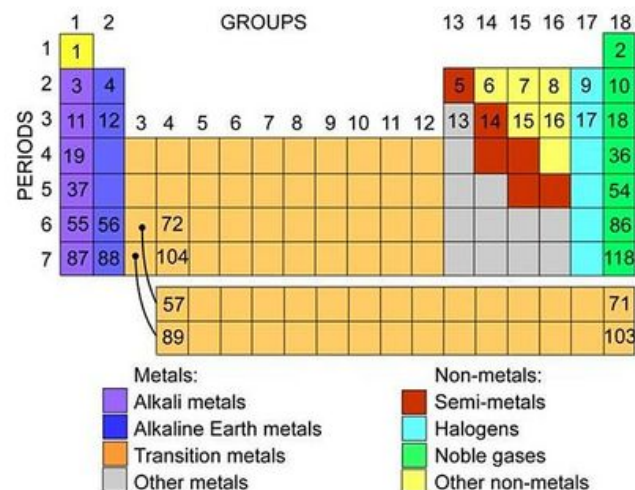
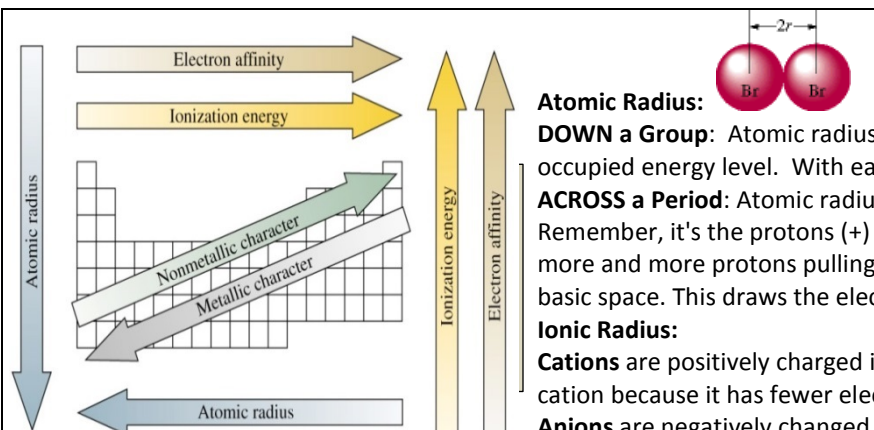| Alkali Group 1	One electron in outer energy level
	Bond with other elements by losing one electron and becoming an ion with a +1 charge (ionic bond)
	Most reactive metals / Francium is the most reactive / Relatively soft metals
Alkaline Group 2	Two electrons in outer energy level
	Bond with other elements by losing two electron and becoming an ion with a +2 charge (ionic bond)
	Form solutions with a pH greater than 7(basic/alkaline)
Halogens Group 17	Seven electrons in outer energy level
	Bond by gaining one electron and becoming a -1 ion (ionic bond)or sharing an electron with another atom(covalent bond)
	Most reactive nonmetals / Fluorine is the most reactive
Noble Gases Group 18	Eight electron in outer energy level
	Generally they do not form bonds / Most Un-reactive (stable)
Transition Metals	Variety of oxidation numbers. Not very reactive

5C Periodic Trends Explanations for PT Trends

Group

increasing nuclear charge

increasing n

1. Orbitals are larger.
2. Electrons are farther from the nucleus.
3. Valence electrons are less tightly bound.

Period

increasing nuclear charge

increasing n

1. Electrostatic attraction increases.
2. Electrons are pulled closer to the nucleus.
3. Valence electrons are more tightly held.

Atomic Radius:

DOWN a Group: Atomic radius INCREASES as you go DOWN a Group because each successive Period (row) has an additional occupied energy level. With each additional energy level the distance between nucleus and valence energy level increases.

ACROSS a Period: Atomic radius DECREASES as you go ACROSS a Period because the net nuclear charge increases (Huh?). Remember, it's the protons (+) in the nucleus that pull on or attract the electrons in the orbitals. Across a Period you are adding more and more protons pulling on electrons occupying the same orbitals. The overall effect is more pulling power in the same basic space. This draws the electrons in closer, making the overall atomic radius smaller as you go left to right across a period.

Ionic Radius:

Cations are positively charged ions formed when metals lose electron(s). Cations are smaller than the metal atom that forms the cation because it has fewer electrons and lose an outer energy level.

Anions are negatively changed atoms formed when a nonmetal gains an electron. The anion is larger than the nonmetal that forms the anion because it has more electrons.

1ˢᵗ Ionization Energy: Amount of energy it takes to remove the most loosely held electron.

DOWN a Group: 1ˢᵗ Ionization energy DECREASES as you go DOWN a Group because the farther the valence electrons are from the nucleus the weaker the attraction between the valence electrons and the nucleus. Less energy it costs for another atom to steal them.

ACROSS a Period: 1ˢᵗ Ionization energy INCREASES as you go ACROSS a Period because the valence electrons are closer to the nucleus resulting in a greater attraction between the valence electrons and the nucleus. The atomic radius is getting smaller, so the protons do a great job of holding on tighter.

Electronegativity: the ability of the atoms of an element to attract the electron(s) of the atoms of other elements

DOWN a Group: Electronegativity DECREASES as you go DOWN a Group because the valence electrons are increasingly farther away from the attraction of the protons in the nucleus. Less pull, less "desire" to grab other electrons.

ACROSS a Period: Electronegativity INCREASES as you go ACROSS a Period because the number of protons (+ charges) in the nucleus increases. More protons in the nucleus means electrons are more strongly attracted to the nucleus.


6A Evidence of Atomic Theory

Dalton's Atomic Theory:

1.   All matter consists of tiny particles called atoms.

2.   Atoms are indestructible and unchangeable. Atoms of an element cannot be created, destroyed, broken into smaller parts or transformed into atoms of another element. Dalton based this hypothesis on the law of conservation of mass and on centuries of experimental evidence.

3.   All atoms of the same element are identical. Elements are characterized by the mass of their atoms. All atoms of the same element have identical weights,

4.   Atoms combine in new ways during a chemical change. When elements react, their atoms combine in simple, whole-number ratios.

Two criteria are usually applied to any theory. First, does it agree with facts which are already known? Second, does it predict new relationships and stimulate additional observation and experimentation? **The Theory has been modified to account for new observations.**

1. Atoms are indestructible and unchangeable. Atoms of an element cannot be created, destroyed, broken into smaller parts or transformed into atoms of another element. We now know elements have subatomic particles and can be transformed by nuclear change but not by physical or chemical change.

2. All atoms of the same element are identical. Elements are characterized by the mass of their atoms. All atoms of the same element have identical weights. Atoms of the same element may have different numbers of neutrons but all atoms of the same elements have the same number of protons. Now an atom is defined as the smallest particle of an element that retains the properties of the element.

J.J. Thompson		
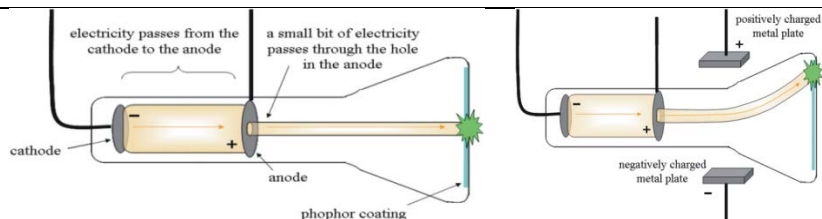
Thompson plum pudding model of the atom — Negative electron plums
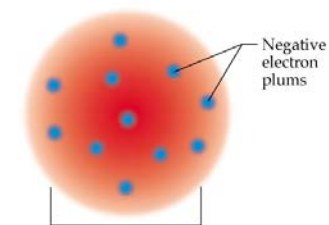
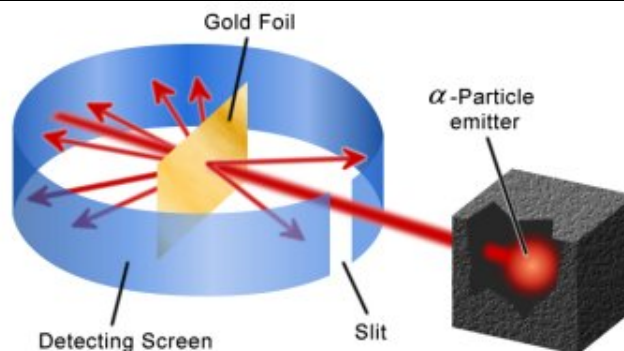| Experiment J.J. Thompson	Observation J.J. Thompson	Conclusion J.J. Thompson
J. J. Thomson painted a substance known as "phosphor" so he could see exactly where the cathode rays hit because the cathode rays made the phosphor glow.	The Cathode ray traveled away from the negative cathode and towards the positive anode.	Cathode Ray may be negatively charged.
He placed a positively charged metal plate on one side of the cathode ray tube and a negatively charged metal plate on the other side of the cathode ray tube, as shown in the above figure.	The flow of the cathode rays passing through the hole in the anode was bent upwards towards the positive metal plate and away from the negative metal plate.	Using the "opposites attract, likes repel" rule, J. J. Thomson argued that if the cathode rays were attracted to the positively charged metal plate and repelled from the negatively charged metal plate, they themselves must have a negative charge!
He put an object in the path of the cathode ray in order to observe the shadow that would result. Energy that travels as waves will bend around the object resulting in a shadow with fuzzy edges. If the ray is a stream of particles the ray will not bend around the object and the shadow will have sharp edges.	The edges of the shadow were sharp (not fuzzy). No particles can be seen. Thomson's experiment also revealed the electron has a very large charge-to-mass ratio.	The cathode ray consists of very small particles.

Based on what was learned about electrons, two other inferences were made about atomic structure.

1. Because atoms are electrically neutral, they must contain a positive charge to balance the negative electrons.
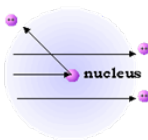
2. Because electrons have so much less mass than atoms, atoms must contain other particles that account for most of their mass

Final Conclusion **Plum Pudding model of the atom**.

Ernest Rutherford		

Experiment Ernest Rutherford	Observations Ernest Rutherford	Conclusion Ernest Rutherford
Gold Leaf Scattering Experiment 1. fire massive alpha particles at the atoms in thin metal foil 2. alpha particles should pass like bullets straight through soft plum pudding atoms	Most of the alpha particle went straight through the gold leaf. But to his great surprise a few alpha particles ricocheted!	**Rutherford's Atom** • concentrated mass and positive charge at the nucleus • electrons roam empty space around the nucleus Most of the alpha particles went straight through the gold leaf because **most of the atom was empty space**. All the mass and positive charged occupies about 1/10,000 the volume. (this is our nucleus). **The electrons are in the empty space outside the center of very dense positive charge called the nucleus**

The nucleus has to be extremely dense for the alpha particles to be deflected. For the nucleus to be dense enough to reflect the alpha particles it had to have very little volume and almost all the mass.

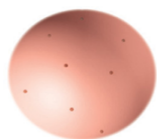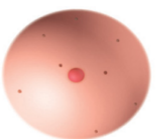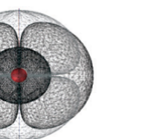## Niels Bohr Observations, Conclusions and Model of the Hydrogen Atom
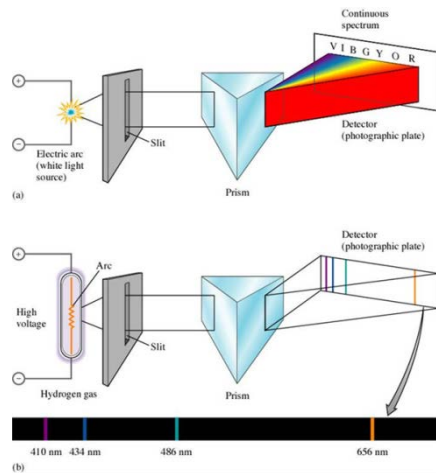
Observation: that light emitted from samples of atom exposed to energy (energizing the atoms) when passed through a prism emit distinct wavelengths of light line spectrum instead of emitting all the colors of light. Each element emits the same line spectrum and are just as characteristic to that element as finger prints are to people.
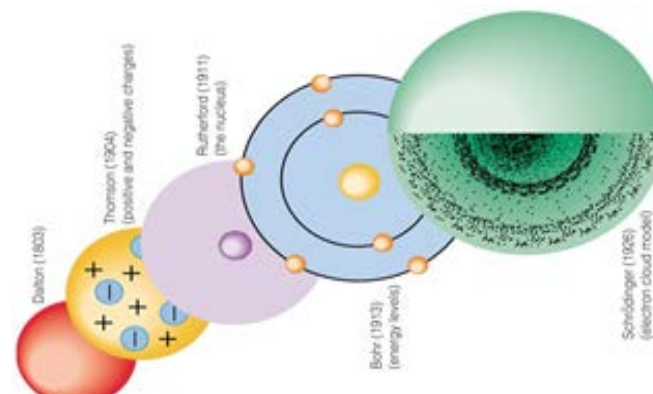
Hypothesis: if energized atoms emit only discrete wavelengths, then electrons can have only discrete energies
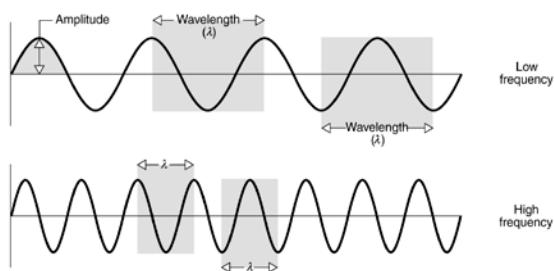
Conclusion:

Bohr's Theory

• Electrons occupy orbits at discrete distances from the nucleus called energy levels.
• Electrons in an non-energized atom occupy the lowest energy orbit (closest to nucleus this is called ground state.)
• Electrons in energized atoms absorb just enough energy to move from a lower energy orbital to a higher energy orbit (further from nucleus) this is called excited state.
• Electrons do not remain in excited state and return to ground state releasing energy that corresponds to all the possible difference in energy of the allowed orbits in the atom.

1803 Dalton proposes the indivisible unit of an element is the atom.

1904 Thomson discovers electrons, believed to reside within a sphere of uniform positive charge (the plum pudding model).

1911 Rutherford demonstrates the existence of a positively charged nucleus that contains nearly all the mass of an atom.

1913 Bohr proposes fixed circular orbits around the nucleus for electrons.

1926 In the current model of the atom, electrons occupy regions of space (orbitals) around the nucleus determined by their energies.

6B Electromagnectic Waves

Electromagnetic radiation can travel through empty space. Most other types of waves must travel through some sort of substance. For example, sound waves need either a gas, solid, or liquid to pass through in order to be heard.

The speed of light is always a constant. c= **(Speed of light: 2.99792458 X 10^8 m s^{-1} or 3.00 X 10^8 m/s) Wavelength** (λ) is the distance of one full cycle of the oscillation. **Wavelengths** are measured between the distances between two consecutive points on a wave which can be between consecutive troughs and crests. It is usually characterized by the Greek symbol ⋏.

Frequency (f) is defined as the number of cycles per second, and is expressed as sec^{-1} or Hertz (Hz).

<u>Frequency is directly proportional to energy (High frequency =High energy) and can be express as: **E =hf**</u>

Where E is energy, h is Planck's constant and f is frequency. Planck's constant: h= 6.62607 x 10^{-34} J s

$$C = \lambda f$$

Where **c** is the speed of light, **λ** stands for the wavelength, and **ν** (nu) stands for frequency. **Shorter wavelength means greater frequency, and greater frequency means higher energy.** Wavelengths are important in that they tell one what type of wave one is dealing with.

In general as the energy increases the wavelength decreases and the frequency increases. $\uparrow E \downarrow \lambda \uparrow f$

The opposite is also true as the energy decreases the wavelength increases and the frequency decreases. $E \uparrow \lambda \downarrow f$
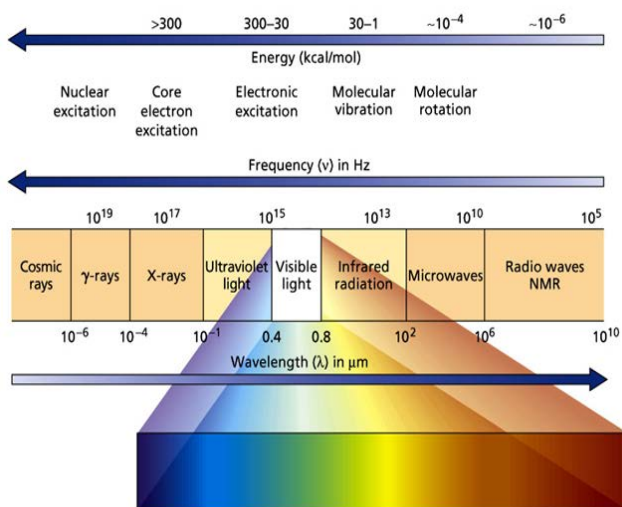
6C Calculating Wavelength, Frequency and Energy

1) A laser used in eye surgery to fuse detached retinas produces radiation with a wave length of 640.0 nm. Calculate the frequency

$$c = \lambda f$$

$$\frac{c}{\lambda} = \frac{\lambda f}{\lambda} \ or \ f = \frac{c}{\lambda}$$

$$f = \frac{3 \times 10^8 \frac{m}{s}}{640.0 \ nm \ \times \frac{1 \ m}{1 \times 10^9 nm}} = \mathbf{4.69 \times 10^{14}} \frac{1}{s} \ or \ hertz$$

2) How much energy is in a photon of the laser light? (using the frequency calculated above)

$$E = hf = 6.626 \times 10^{-34} Js \times 4.69 \times 10^{14}\, 1/s = \textbf{3.11 x 10}^{\textbf{-19}}\, \textbf{J per photon}$$

3) Visible light has frequency between about 4.0 x 10^{14} hertz and 7.9 x 10^{14} hertz or 1/s. What is the wavelenght of the lowest frequency of light?

$$c = \lambda f \quad \lambda = \frac{c}{f} =$$

$$\frac{3 \times 10^{8} m/s}{4.0 \times 10^{14}\, 1/s}$$

$$= \textbf{7.5 x 10}^{\textbf{-7}}\, \textbf{m}$$

6D Calculating Average Atomic Mass

$$\text{Relative abundance} = \frac{\%\ abundance}{100}$$

(atomic mass of isotope × relative abundance) × (atomic mass of isotope × relative abundance) ⋯ = average atomic mass

Example:

The most abundant isotopes of silicon are silicon-28, silicon-29, and silicon-30. Given the atomic masses and abundances in the table below, what is the average atomic mass of silicon?

Isotopic Composition of Silicon		
Isotope	Atomic mass (amu)	Approximate abundance (percent)
Silicon-28	27.97693	92.23
Silicon-29	28.97649	4.68
Silicon-30	29.97377	3.09

Remember we are changing percent abundance to relative abundance by dividing by 100 or moving the decimal two places to the left. The equation above will be used to calculate the atomic mass in amu.

$(27.97693\ amu \times .9223) + (28.97649\ amu \times .0468) + (29.97377\ amu \times .0309) = \textbf{28.1 amu}$

6E Electron Configurations & Lewis Dot Structures

Atomic Orbitals and Electrons in Principal Energy Levels			
Principal energy level	Type of sublevel	Number of orbitals in sublevels	Maximum number of electrons
1	s	1	2
2	s, p	1 + 3 = 4	8
3	s, p, d	1 + 3 +5 = 9	18
4	s, p, d, f	1 + 3 + 5 + 7 = 16	32

Principal quantum numbers describe the size, shape, and energy of the orbital an orbital is probable location of an electron.

- Principal energy level (principal quantum number)=n describes the size
- Sublevel = ℓ describes the shape of the probable location of an electron
- Orbital = each sublevel has a set number of orbitals of a given size and shape that can be occupied by a maximum of 2 electrons.

The location closest to the nucleus has the least amount of energy. This energy increases as you get further from the nucleus. Electrons occupy the lowest energy orbitals first.

Energy level	Number of sublevels	Names of sublevels
1	1	s
2	2	s, p
3	3	s, p, d
4	4	s, p d, f
5	5	s, p, d, f, g

Sublevel Name	s	p	d	f
Number of Orbitals	1	3	5	7
Maximum number of electrons	2	6	10	14

Electron Configurations: Assigning each electron in an atom to the energy level and sublevel it occupies in the atom.   Number of Electrons

Energy Level ⟵⟶ $1s^2$ ⟵⟶ Sublevel

Each orbital can hold no more than two electrons. Two electrons in the same orbital must have

Example: Ni atomic number is 28 so it has 28 protons and 28 electrons because it does not have a charge. Electron 1 and 2 are in 1s, the 3rd and 4th electrons are in 2s, the next six electrons which are the 5th to the 10th are in the 2p, the 11th and 12th electron are in the 3s, the next six electrons which are 13th to the 18th are in the 3p sublevels orbitals, the 19th and 20th electron occupy the 4s orbital, the last 8 electrons will occupy the 3d orbitals but not fill all the orbitals in the 3d sublevel. **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$**

Yes there is a short cut: the short cut is sometimes called core notation:
Using Ni as an example:
Use the nearest noble gas with a smaller atomic number than the element you are writing the electron core notation for: for Ni this would be Ar with the atomic number 18. Write the symbol of the noble gas in [ ] with the number of electron of the noble gas as a superscript. $[Ar]^{18}$ (this is saying the first 18 electrons in Ni are in the same orbitals as the 18 electrons in Ar. Add the rest of the electrons reading from the periodic table. You will be assigning the 19th through 28th electron. Remember the 19th electron is in the same orbital for all atoms.
$[Ar]^{18} 4s^2 3d^8$ or $[Ar]^{18} 3d^8 4s^2$

s block

1 1A	2 2A
3	4
11	12
19	20
37	38
55	56
87	88

d block

3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B
21	22	23	24	25	26	27	28	29	30
39	40	41	42	43	44	45	46	47	48
71	72	73	74	75	76	77	78	79	80
103	104	105	106	107	108	109	110	111	112

p block

13 3A	14 4A	15 5A	16 6A	17 7A	2
5	6	7	8	9	10
13	14	15	16	17	18
31	32	33	34	35	36
49	50	51	52	53	54
81	82	83	84	85	86
113	114	115	116	117	118

1s: 1
18 8A

f block

4f	57	58	59	60	61	62	63	64	65	66	67	68	69	70
5f	89	90	91	92	93	94	95	96	97	98	99	100	101	102

Lewis Dot Diagrams

- A Lewis dot structure is like a simplified electron energy level model.
- The Lewis structure contains the element symbol with dots representing electrons.
- The only electrons shown are those on the outer energy level or valence electrons.
- The electrons are placed around the element symbol, one at a time, clockwise or counterclockwise, and then grouped in pairs as more electrons are added.

PERIODIC TABLE ELEMENTS 1-20

HYDROGEN 1 H·							HELIUM 2 He·
LITHIUM 3 Li·	BERYLLIUM 4 Be·	BORON 5 ·B·	CARBON 6 ·C·	NITROGEN 7 ·N·	OXYGEN 8 ·O·	FLOURINE 9 ·F·	NEON 10 ·Ne·
SODIUM 11 Na·	MAGNESIUM 12 Mg·	ALUMINUM 13 ·Al·	SILICON 14 ·Si·	PHOSPHORUS 15 ·P·	SULFUR 16 ·S·	CHLORINE 17 ·Cl·	ARGON 18 ·Ar·
POTASSIUM 19 K·	CALCIUM 20 Ca·						

Hydrogen $1s^1$ H·

Oxygen $1s^2 2s^2 2p^4$ ·Ö·

Chlorine $1s^2 2s^2 2p^6 3s^2 3p^5$:Cl·

Chloride ion Cl^{-1} $1s^2 2s^2 2p^6 3s^2 3p^6$:Cl:

:F· + e⁻ ⟶ [:F:]⁻

·Na ⟶ [Na]⁺ + e⁻
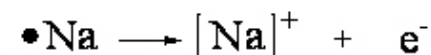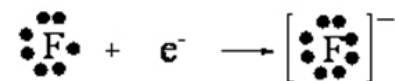
7A Naming Compounds

Must Memorize to Name Covalent Compounds

Prefixes used for naming covalent compounds	
Number	**Prefix**
1	mono- *
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

*The prefix mono- is only used for the more electronegative element.

Need to know the following to name acids

Anion ending	Formula	Acid name	Example
ide	HCl	hydro-(stem)-ic acid	hydrochloric acid
-ite	H_2SO_3	(stem)-ous acid	sulfurous acid
-ate	HNO_3	(stem)-ic acid	nitric acid

Need to know when to add roman numerals to metal name

Ion Name	Ion	Ion Name	Ion
Iron(II)	Fe^{2+}	Silver(I)	Ag^+
Iron(III)	Fe^{3+}	Silver(II)	Ag^{2+}
Copper(I)	Cu^+	Mercury(I)	Hg^+
Copper(II)	Cu^{2+}	Mercury(II)	Hg^{2+}

Three Types of Compounds to Name:
1. Ionic compounds: 1st element a metal
2. Covalently bonded compounds: 1st element a nonmetal
3. Acid: 1st element is H

POLYATOMIC IONS

Acetate	$C_2H_3O_2^-$, CH_3COO^-
Ammonium	NH_4^+
Carbonate	CO_3^{2-}
Chlorate	ClO_3^-
Chlorite	ClO_2^-
Chromate	CrO_4^{2-}
Cyanide	CN^-
Dichromate	$Cr_2O_7^{2-}$
Hydrogen carbonate	HCO_3^-
Hydroxide	OH^-
Hypochlorite	ClO^-
Nitrate	NO_3^-
Nitrite	NO_2^-
Perchlorate	ClO_4^-
Permanganate	MnO_4^-
Phosphate	PO_4^{3-}
Sulfate	SO_4^{2-}
Sulfite	SO_3^{2-}

Reference Chart Provided On Test

Naming Ionic Compounds: First Element is a Metal

First word
Ionic name of metal
(roman numerial is added only if the metal has more than one possible charge)

+

Second word
name of anion(if it is an ion of one nonmetal or OH⁻ or CN⁻ the ion ends in ide)
polyatomic ions other than OH⁻ and CN- can end in ate or ite

KNO_3 is named potassium nitrate.
Iron in FeO must have a +2 charge because oxygen has a-2 charge (O^{-2}). The compound's name is iron(II) oxide
Tin in SnS_2 must have a +4 charge because 2 S^{-2} have a -4 charge. The name of the compound is tin(IV) sulfide
$BaCl_2$ is named barium chloride
Na_2SO_3 is named sodium sulfite
$K_2Cr_2O_7$ is named potassium dichromate
Cu_3PO_4 . PO_4 has a -3 charge so each of the 3 Cu must be +1 to = +3. The name of the compound is copper(I) phosphate

Naming Covalently Bonded Compounds: First Element a Nonmetal

| First word
*Prefix + name of first element | + | Second word
*Prefix + usually first sylabol of second element + ide |

*The prefix mono- is only used for the more electronegative element. Prefixes are listed in a table on the page above.

CO_2 Carbon dioxide As_2Cl_5 diarsenic pentachloride CCl_4 carbon tetrachloride

Naming Acids: First Element is H: Two Type of Acids Binary and Nonbinary
Binary Acid contains only 2 elements

| First word
hydro-(stem)-ic   (the stem is usually the first syllable of second element) | + | Second word
acid |

HBr(aq) = hydrobromic acid H_2S(aq) = hydrosulfuric acid

Nonbinary Acid: Hydrogen + Polyatomic Ion

| First word
If the acid ends in polyatomic ion with sufix ite the ite becomes ous
If the acid ends in polyatomic ion with sufix ate the ate becomes ic | + | Second word
acid |

$HClO_3$ Chloric acid $HClO_2$ chlorous acid H_2SO_4 sulfuric acid H_2SO_3 sulfrous acid

---

7B Writing Formulas
All compounds other than covalent
Cross over the combining capacities from each of the ions so that they become the subscripts of the opposite ions. Reduce the subscripts to **lowest terms** if necessary.
Subscripts of 1 are not written.
Remember:
1. Write positive ion first- write negative ion which maybe a monoatomic or polyatomic ion (more than one element)
2. Polyatomic do not change formula or subscripts- when adding a subscript ion is enclosed in parenthesizes.
2. Determine the charge on each ion
3. Crisis cross and reduce

Write formula for hydrosulfuric acid

H^{+1} S^{-2}

H_2 S no need to reduce = **H_2S**
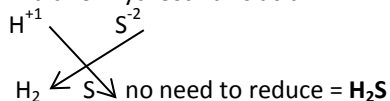
Write the formula for tin(IV) sulfate

Sn^{+4} SO_4^{-2}

Sn_2 $(SO_4)_4$ reduce = **$Sn(SO_4)_2$**
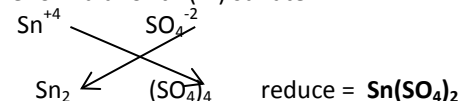
Binary Covalent Compounds
The prefix on eaxh word of the name indicates the number of eachelement in the compound
Write the elements in the same order as they are in the name
Write the formula fo diphosphorous pentaoxide → P_2O_5

Rules for Drawing Lewis Dot Diagrams of Covalent Compounds:

1. Calculate the number of valence electrons for each atom and the total for the molecule or ion. For ions, add or subtract electrons as necessary to obtain the correct charge.
2. Write the skeletal structure; place correct number of electrons around each atom.
3. Replace pairs of electrons with a bond
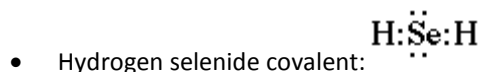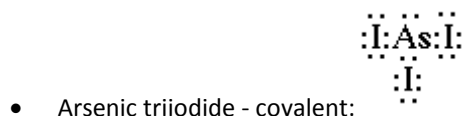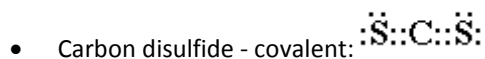4. Make certain each atom has a complete octet. Use multiple bonds if necessary to complete octet. Double check total number of electrons.

Additional Hints:

H is always an end (terminal) atom. It has one electron and thus is only able to bond covalently with one other atom.
Halogens and Oxygen are often terminal atoms.
In binary compounds, the central atom has the lowest subscript and is usually listed first.
Carbon has four bonds in most compounds.

- Barium iodide- ionic: $[\ddot{\ddot{I}}:]^- \; Ba^{+2} \; [\ddot{\ddot{I}}:]^-$

- Carbon disulfide - covalent: $:\ddot{S}::C::\ddot{S}:$

- Arsenic triiodide - covalent: $:\ddot{I}:\overset{\cdot\cdot}{As}:\ddot{I}:$ $\;\;\;:\ddot{I}:$

- Hydrogen selenide covalent: $H:\overset{\cdot\cdot}{Se}:H$
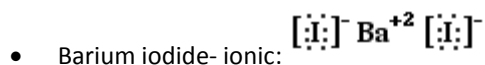
7D Metallic Bonds Explain Properties of Metals

Metallic Bonding

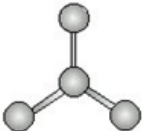Free Electrons from outer shell of metal atoms
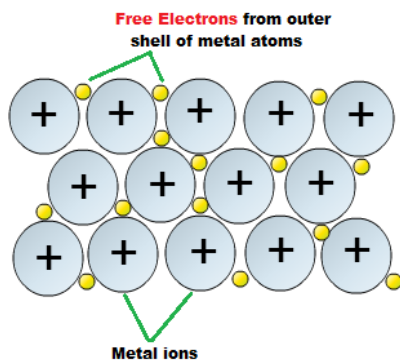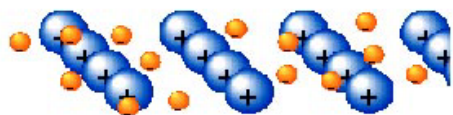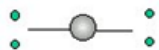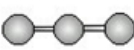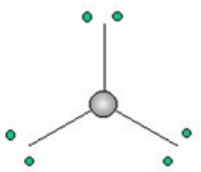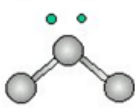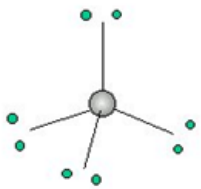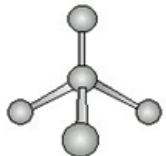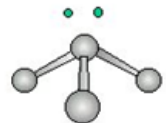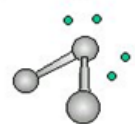
Metal ions

Key Concepts
1. A metal is a lattice of positive metal 'ions' in a 'sea' of delocalized electrons.
2. Metallic bonding refers to the interaction between the delocalized electrons and the metal nuclei.
3. The physical properties of metals are the result of the delocalization of the electrons involved in metallic bonding

Examples:
1. Melting Points general, metals have high melting and boiling points because of the strength of the metallic bond. The strength of the metallic bond depends on the number of electrons in the delocalized 'sea' of electrons. (More delocalized electrons results in a stronger bond and a higher melting point.)
2. The delocalized electrons are free to move in the solid lattice. These mobile electrons can act as charge carriers in the conduction of electricity or as energy conductors in the conduction of heat
3. The delocalized electrons are free to move in the solid lattice. These mobile electrons can act as charge carriers in the conduction of electricity or as energy conductors in the conduction of heat
4. Metals typically have a shiny, metallic luster. Photons of light do not penetrate very far into the surface of a metal and are typically reflected, or bounced off, the metallic surface.
5. Metals are malleable and ductile.
The delocalized electrons in the 'sea' of electrons in the metallic bond, enable the metal atoms to roll over each other when a stress is applied.

Total Electron Pairs	Electron Pair Geometry	Bonding Pairs	Nonbonding Pairs	Molecular Geometry
2	Linear	2	0	Linear
3	Trigonal planar	3	0	Trigonal planar
		2	1	Bent

4	Tetrahedral	4	0	Tetrahedral
		3	1	Trigonal pyramidal
		2	2	Bent

8A Define & Use Moles

The mole (mol) is a unit for counting objects. It is especially useful for counting tiny objects like atoms, molecules, ions, and formula-units.

1 mole of **particles** (atoms, ions or molecules) = 6.02 x 10^{23} particles for any substance!

Other counting Units are:

A ream of paper contains 500 sheets of paper.

A dozen eggs contains 12 eggs.

- ✓ 6.02 x 10^{23} atoms (particles) = the atomic mass in grams
- ✓ 6.02 x 10^{23} atoms (particles)= 1 mole
- ✓ 1 mole = the atomic mass in grams

How many atoms make up 3.5 mol of Ag?

$$\frac{3.5 \ \cancel{mol \ Ag}}{1} \times \frac{6.02 \ \times \ 10^{23} \text{atoms Ag}}{1 \ \cancel{mol \ Ag}} = \mathbf{2.11} \ \times \mathbf{10^{24} \ atoms \ Ag}$$
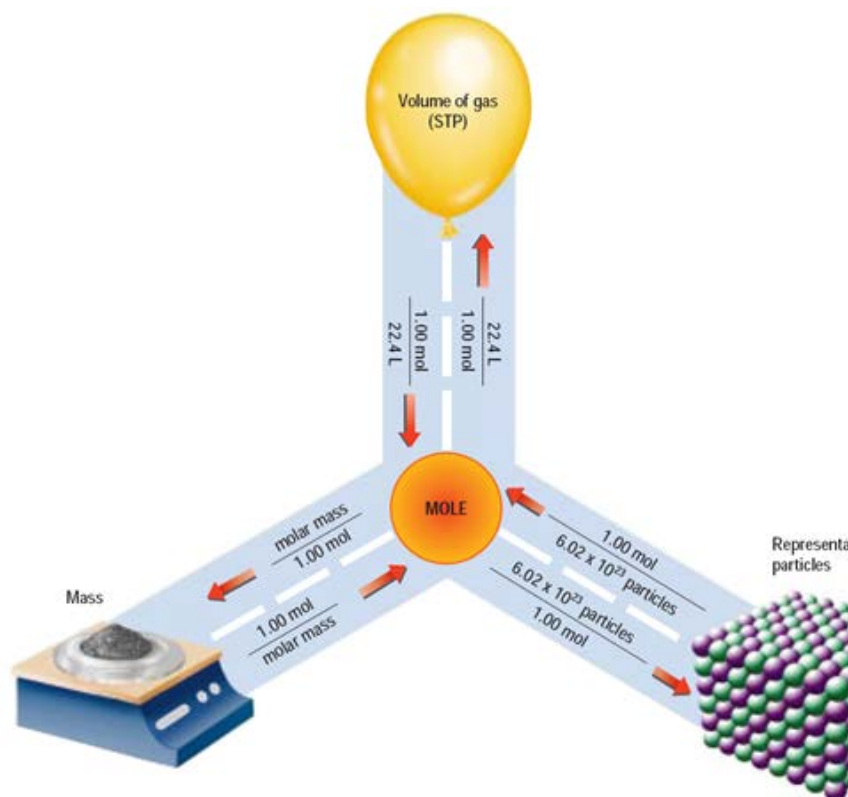
How many molecules are there in 5.5 moles of carbon dioxide?

$$\frac{5.5 \ \cancel{mol \ CO_2}}{1} \times \frac{6.02 \ \times \ 10^{23} \text{ molecules } CO_2}{1 \ \cancel{mol \ CO_2}} = \mathbf{3.31} \ \times \mathbf{10^{24} \ molecules \ CO_2}$$

How many moles are 7.0 X 10^{15} atoms of sulfur?

$$\frac{7.0 \ \times \ 10^{15} \cancel{\text{atoms of S}}}{1} \times \frac{1 \text{ mole of S}}{6.02 \ \times \ 10^{23} \ \cancel{\text{atoms of S}}} = \mathbf{1.16} \ \times \ \mathbf{10^{-8} \ mole \ S}$$

8B Calculating Atoms, Ions, Molecules Using Moles

How to Determine the Molar Mass of a Molecule:

To find the molar mass of a compound:
- Use the chemical formula to determine the number of each type of atom present in the compound.
- Multiply the atomic weight (from the periodic table) of each element by the number of atoms of that element present in the compound.
- Add it all together

Molar Mass of $Zn(NO_3)_2$:

In this compound, we have 1 atom of zinc, 2 atoms of nitrogen ($1 \times 2 = 2$) and 6 atoms of oxygen ($3 \times 2 = 6$). The molar mass of zinc nitrate =

(1 atom Zn $\times$ 65 g of zinc) = 65 g Zn

(2 atoms N $\times$ 14 g of nitrogen) = 28 g N

(6 atoms O $\times$ 16 g of oxygen) = 96 g O

65 + 28 + 96 = **189 g $Zn(NO_3)_2$**

> ⁕ 6.02×10^{23} atoms (particles) = the atomic mass in grams
> ⁕ 6.02×10^{23} atoms (particles)= 1 mole
> ⁕ 1 mole = the atomic mass in grams

Example 1: How many moles are present in 34 grams of $Cu(OH)_2$?

Use conversion between moles and grams

1 mole = molar mass

determine molar mass = (1 mole Cu $\times$ 63.55 g Cu) + (2 mole O $\times$ 16.00 g O) + (2 mole H $\times$ 1.00 g H) = 97.55 g $Cu(OH)_2$

$$\frac{34 \text{ g } \cancel{Cu(OH)_2}}{1} \times \frac{1 \text{ mole } Cu(OH)_2}{97.55 \text{ g } \cancel{Cu(OH)_2}} = \mathbf{0.35 \text{ mole } Cu(OH)_2}$$

Example 2: How many atoms of aluminum are there in a 7.85 g sample of aluminum? (molar mass of Al is ? 26.982 g

$$\frac{7.85 \text{ } \cancel{\text{g Al}}}{1} \times \frac{1 \text{ } \cancel{\text{mol Al}}}{16.982 \text{ } \cancel{\text{g Al}}} \times \frac{6.02 \times 10^{23} \text{ Al atoms}}{1 \text{ } \cancel{\text{mol Al}}} = \mathbf{1.75 \times 10^{23} \text{ Al atoms}}$$

Example 3: 16.2 g of magnesium nitride contains how many magnesium ions?

write the formula for magnesium nitride →Mg_3N_2

- restate question 16.2 g Mg_3N_2 = _____ Mg^{+3} ions
- Identify conversions necessary to answer question. You can start by writing conversion you know and deciding if they help answer the question.
- 1 mole Mg_3N_2 = molar mass of g Mg_3N_2
- Calculate molar mass (You should get 100.929 g Mg_3N_2)
- How many Mg^{+2} ions are in 1 mole of Mg_3N_2??? 1mole Mg_3N_2 = 3 mole of Mg^{+2}
- 1 mole Mg^{+2} = 6.02×10^{23} mol Mg^{+2}

$$\frac{16.2 \text{ g } Mg_3N_2}{1} \times \frac{1 \text{ mol } Mg_3N_2}{100.929 \text{ g } Mg_3N_2} \times \frac{3 \text{ mol } Mg^{+2}}{1 \text{ mol } Mg_3N_2} \times \frac{6.02 \times 10^{23} Mg^{+2} \text{ions}}{1 \text{ mol } Mg^{+2}} = \mathbf{2.90 \times 10^{23} \ Mg^{+2} \ ions}$$

Example 4: 12.8 of dinitrogen tetraoxide contains howmany molecules of dinitrogen tetroxide?

$$\frac{12.8 \text{ g } N_2O_4}{1} \times \frac{1 \text{ mol } N_2O_4}{92.010 \text{ g } N_2O_4} \times \frac{6.02 \times 10^{23} \text{molecules } N_2O_4}{1 \text{ mol } N_2O_4} = \mathbf{8.37 \times 10^{23} \ molecules \ of \ N_2O_4}$$

8C Calculating Percent Composition, Empirical Formulas and Molecular Formulas

Percent compostion is the % be mass of each element in a compound

$$percent \ of \ an \ element \ in \ a \ compound = \frac{molar \ mass \ of \ the \ element}{molar \ mass \ of \ compound} \times 100$$

Example 1: Calculate the percent of magnesium in $Mg(NO_3)_2$.

Step 1

1Mg = 1 x 24 = 24
2N = 2 x 14 = 28
6O = 6 x 16 = 96
24 + 28+ 96 = 148g/mole

Step 2 =

$$\frac{24 \text{ g Mg}}{148 \text{ g } Mg(NO_3)_2} \times 100 = \mathbf{16.2 \ \% \ Mg}$$

Molecular Formula	Empirical Formula
H_2O	H_2O
CH_3COOH	CH_2O
CH_2O	CH_2O
$C_6H_{12}O_6$	CH_2O

Empirical Formula - Simplest formula for a compound; it is the smallest whole-number ratio of the atoms present

Molecular Formula - Expression of the formula for a compound; it shows the actual number of atoms of each element present in one molecule of the compound.

Empirical Formula: Simplest formula for a compound; it is the smallest whole-number ratio of the atoms present.

Here's a rhyme:

Percent or data to mass
Mass to mole
Divide by smallest
Multiply 'til whole

Example 2: Calculate the empirical formula for a compound composed of 26.6% potassium, 35.4% chromium, and 38.1% oxygen.

Percent or data to mass: First, assume that you are given a 100-g sample, which would mean that the mass of each element in grams is equal to the percent composition value. If given data in grams you will need to calculate grams of each element using the data.

From grams we find moles of each element in the formula.

Mass to mole

Find **moles potassium**.

$$\frac{26.6 \text{ g K}}{1} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} = 0.6803 \text{ mol K}$$

Next find moles of chromium.

$$\frac{35.4 \text{ g Cr}}{1} \times \frac{1 \text{mol Cr}}{52.00 \text{ g Cr}} = 0.6808 \text{ mol Cr}$$

Next find moles of oxygen.

$$\frac{381 \text{ g O}}{1} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.381 \text{ mol O}$$

Divide by smallest Next, set up a mole ratio that relates the moles of each element to the moles of the element that is least present.

$$K = \frac{0.6803}{0.6803} = 1 \quad Cr = \frac{0.6808}{0.6803} = 1.0007 = 1 \quad O = \frac{2.381}{0.6803} = 3.4999 = 3.5$$

Multiply 'til whole Because a chemical formula must have only whole numbers we multiply by 2 to yield:

K = 1 × 2= 2   Cr = 1 × 2= 2   O = 3.5 × 2= **7**   **These numbers are the subscripts $K_2Cr_2O_7$** as the empirical formula.

To determine the Molecular Formula you need to be given:
1. Empirical formula
2. Molar mass of compound

Molecular Formula:
1. Calculate empirical formula if not given
2. Divide the molar mass by the empirical formula mass.
3. Multiply all the subscripts in the empirical formula by the number you get when you divide the molar mass by the mass of the empirical formula

$C_6H_{12}O_6$

If subscripts in a formula will reduce, it is NOT an empirical formula!
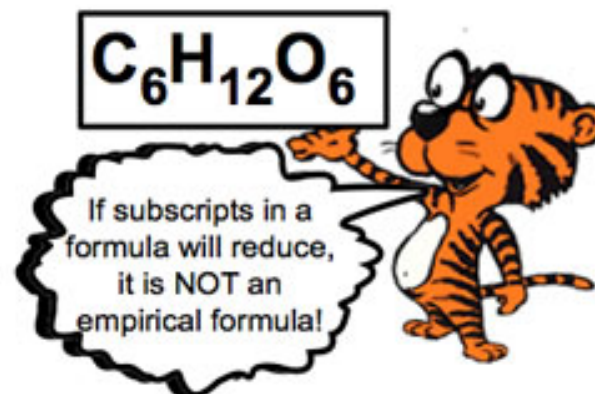
Example 3: A 1.000 g sample of red phosphorus powder was burned in air and reacted with oxygen gas to give 2.291 g of a phosphorus oxide. Calculate the empirical formula and molecular formula of the phosphorus oxide given the molar mass is approximately 284 g/mol.

1. Calculate empirical formula:

~~Percent or~~ **data to mass** % not given use data to determine the mass of P and O:

given that P = 1.00g and P_xO_x = 2.291 g . You know the mass of O = mass of P_xO_x – mass of P

$$2.291 \text{ g } P_xO_x - 1.000 \text{ g P} = 1.291 \text{ g O}$$

Mass to mole:

$$\frac{1.00 \cancel{\text{ g P}}}{1} \times \frac{1 \text{ mol P}}{30.97 \cancel{\text{ g P}}} = 0.032289 \text{ mol P}$$

$$\frac{1.291 \cancel{\text{ g O}}}{1} \times \frac{1 \text{ mol O}}{16.00 \cancel{\text{ g O}}} = 0.0806875 \text{ mol O}$$

Divide by smallest:

$$\frac{0.032289 \text{ mol P}}{0.032289 \text{ mol P}} = 1$$

$$\frac{0.0806875 \text{ mol O}}{0.0806875 \text{ mol O}} = 2.48989 \approx 2.50 \text{ mol} - \text{ not close enough to round}$$

Multiply 'til whole

2 x 2.500 is 5 which is whole so multiply each by 2

1 P x 2 = **2 P**   2.50 O x 2 = **5 O**   **Empirical formula is P_2O_5**

2. Divide the molar mass by the empirical formula mass:

Empirical mass (5 × 30.97g P) + (2 x 16.00g O) = 141.943g empirical formula mass

The "empirical formula weight" = 141.943g

$$\frac{284}{142} = 2$$

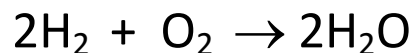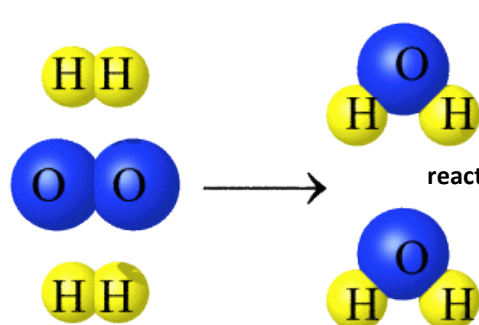3. The molecular formula is a multiple of 2 times the empirical formula:

P= 2 x 2 = 4

O= 2 x 5=10

molecular formula is P_4O_{10}

8D Write Balance Chemical Equations

$$2H_2 + O_2 \rightarrow 2H_2O$$

Equations must be balanced because: Law of Conservation of Matter: Atoms can be neither created nor destroyed in an ordinary chemical reaction, so there must be the same number of atoms on both sides of the equation. **The mass of all the reactants (the substances going into a reaction) must equal the mass of the products (the substances produced by the reaction).**

Two types of numbers are found in a chemical equation:

Subscripts: H_2O

The small numbers to the lower right of chemical symbols. Subscripts represent the number of atoms of each element in the molecule.

Coefficients: $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(g)$
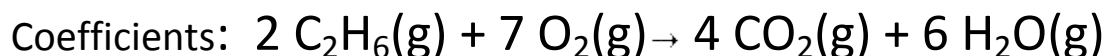
The large numbers in front of chemical formulas. Coefficients represent the number of molecules of the substance in the reaction.

Steps in Balancing a Chemical Equation An Equation is balanced when there are the same numbers of atoms of each element on both sides of the equation.

Balancing a chemical equation involves a certain amount of trial and error. In general, however, you should follow these steps:
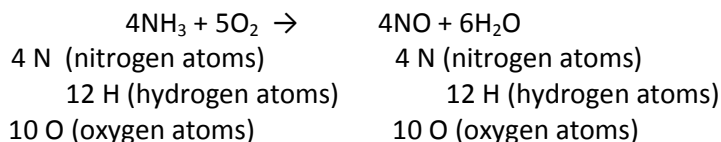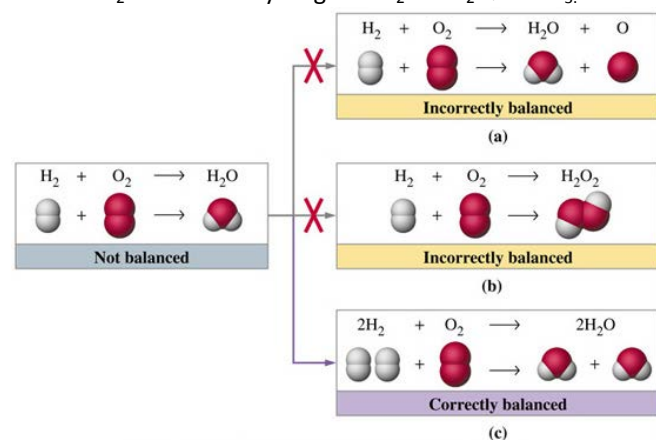
1. Count each type of atom in reactants and products. Does the same number of each atom appear on both sides of the arrow? If not, the equation is not balanced, and you need to go to step 2.
2. Place coefficients, as needed, in front of the symbols or formulas to increase the number of atoms or molecules of the substances. Use the smallest coefficients possible. Warning! Never change the subscripts in chemical formulas. Changing subscripts changes the substances involved in the reaction. Change only the coefficients.
3. Repeat steps 1 and 2 until the equation is balanced.

Example: Balance this chemical equation for the reaction in which nitrogen (N_2) and hydrogen (H_2) combine to form ammonia (NH_3): $N_2 + H_2 \rightarrow NH_3$

1: First count the nitrogen atoms on both sides of the arrow. There are two nitrogen atoms in the reactants so there must be two in the products as well.

Place the coefficient 2 in front of NH_3 to balance nitrogen: $N_2 + H_2 \rightarrow$ **2**NH_3

Now count the hydrogen atoms on both sides of the arrow. There are six hydrogen atoms in the products so there must also be six in the reactants. Place the coefficient 3 in front of H_2 to balance hydrogen: $N_2 +$ **3** $H_2 \rightarrow 2\ NH_3$.

$4NH_3 + 5O_2 \rightarrow$	$4NO + 6H_2O$
4 N (nitrogen atoms)	4 N (nitrogen atoms)
12 H (hydrogen atoms)	12 H (hydrogen atoms)
10 O (oxygen atoms)	10 O (oxygen atoms)

8E
Stoichiometry

Mole X

Mole Y

$$\frac{\text{coefficient moles } Y}{\text{coefficient moles } X}$$

BASED UPON MOLE - MOLE RATIOS.

Basic Pattern:

1. If quantity given is grams Convert to moles
2. Perform mole $\leftrightarrow$ mole conversion
3. Convert to desired units

Stoichiometry Example: If iron pyrite, FeS_2, is not removed from coal, oxygen from the air will combine with both the iron and the sulfur as coal burns. If a furnace burns an amount of coal containing 100 g of FeS_2, how much SO_2 (an air pollutant) is produced? The following example will discuss each conversion. The math is done when all the conversions are done

1. Write a balanced equation showing the formation of iron (III) oxide and sulfur dioxide.

$$4FeS_2 + 11O_2 \longrightarrow 2\ Fe_2O_3 + 8\ SO_2$$
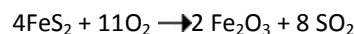
The equation provides conversion factor to convert from moles of FeS_2 to moles of SO_2

4 moles of FeS_2 = 8 moles of SO_2 or 1 mole FeS_2 = 2 mole SO_2
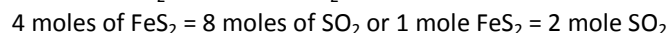
2. Write the mass information given in the problem.

$$\frac{100\ \text{g FeS}_2}{1}$$

3. Convert grams of FeS_2 to moles of FeS_2.

$$\frac{100\ \text{g FeS}_2}{1} \times \frac{1\ \text{mol FeS}_2}{120\ \text{g FeS}_2}$$

4. Changes moles of FeS_2 (reactant) to moles of SO_2 (product).

This ratio comes from the coefficients in the balanced equation.

Notice that the ratio could be reduced from 8 : 4 to 2 : 1 when placed in the factor-label form. Reducing is not necessary.

$$\frac{100\ \text{g FeS}_2}{1} \times \frac{1\ \text{mol FeS}_2}{120\ \text{g FeS}_2} \times \frac{8\ mol\ SO_2}{4\ mol\ FeS_2}$$

5. Convert moles of SO_2 to grams of SO_2

$$\frac{100\ \cancel{\text{g FeS}_2}}{1} \times \frac{1\ \cancel{\text{mol FeS}_2}}{120\ \cancel{\text{g FeS}_2}} \times \frac{8\ \cancel{\text{mol SO}_2}}{4\ \cancel{\text{mol FeS}_2}} \times \frac{64\ \text{g SO}_2}{1\ \cancel{\text{mol SO}_2}} = \textbf{107 g SO}_2$$

LIMITING REAGENT: is the reagent that is totally consumed during the reaction and therefore determines or limits the amount of product formed.

EXCESS REAGENT: is the reactant that is not totally consumed during the reaction and therefore some remains when the reaction has stopped.

THEORETICAL YIELD: is the calculated amount of product that could be produced if no error occurred. Error always occurs so the theoretical yield never equals the measured amount of product produced.

Limiting Problem: A 50.6 g sample of magnesium hydroxide is reacted with 45.0 g of hydrochloric acid according to the reaction:

$Mg(OH)_2$ + 2 HCl $\rightarrow$ $MgCl_2$ + 2 H_2O what is the theoretical yield of $MgCl_2$?

Is this a limiting reagent problem? One way to find out is to write down what is known about any component of the reaction below that component:

Notice how quantities of both reactants are known. Which one will be used up first? You can't tell, nor should you jump to any conclusions. Just because it looks like there is less $Mg(OH)_2$ present does not automatically mean it will be used up before all of the HCl is consumed. This is a limiting reagent problem.

To determine which reactant is the limiting reagent:
1. Calculate the amount of product using the amount of each reactant.
2. The reactant that yield the least amount of product is the limiting reagent.
3. The smallest amount of product calculated is the theoretical yield of this product.

$$\frac{50.6 \text{ g } Mg(OH)_2}{1} \times \frac{1 \text{ mol } Mg(OH)_2}{58.3 \text{ g } Mg(OH)_2} \times \frac{1 \text{ mol } MgCl_2}{1 \text{ mol } Mg(OH)_2} \times \frac{95.3 \text{ g } MgCl_2}{1 \text{ mol } MgCl_2} = 82.7 \text{g } MgCl_2$$

$$\frac{45 \text{ g } HCl}{1} \times \frac{1 \text{ mol } HCl}{36.5 \text{ g } HCl} \times \frac{1 \text{ mol } MgCl_2}{2 \text{ mol } HCl} \times \frac{95.3 \text{ g } MgCl_2}{1 \text{ mol } MgCl_2} = 58.7 \text{ g } MgCl_2$$

HCl is limiting reagent because it produced the least amount of product because it was consumed first. The HCl will be run out before the magnesium hydroxide and thereby limit the amount of product formed. For this reason, use the moles of HCl to calculate the theoretical yield of magnesium chloride:

$$\frac{1.23 \text{ mol } HCl}{1} \times \frac{1 \text{ mol } Mg(OH)_2}{2 \text{ mol } HCl} \times \frac{95.3 \text{g } Mg(OH)_2}{1 \text{ mol } Mg(OH)_2} = \mathbf{58.6 \text{ g } Mg(OH)_2}$$

PERCENT YIELD PROBLEMS:
- Theoretical yield – this is what you calculate from stoichiometry - only in a perfect world
- Actual yield – what you get in a lab in the real world (no yield will be 100% - useful in lab reports)

$$\textbf{Percent yield} = \frac{\text{Actual Yield} \left(\text{measured amount of product produced}\right)}{\text{Theoretical Yield} \left(\text{calculated possible amount of product produced}\right)} \times 100$$

Percent Yield Example:
Consider the combustion of 215.0g of butane, C4H1o with plenty of oxygen. If you are able to collect 295.3g of water in the laboratory, what was your percent yield?

$$\frac{215.0 \text{ g } C_4H_{10}}{1} \times \frac{1 \text{ mol } C_4H_{10}}{59.084 \text{ g } C_4H_{10}} \times \frac{10 \text{ mol } H_2O}{2 \text{ mol } C_4H_{10}} \times \frac{10.015 \text{ g } H_2O}{1 \text{ mol } H_2O} = 333.4 \text{ g } H_2O$$

$$\frac{295.3 \text{ g } H_2O}{333.4 \text{ g } H_2O} \times 100 = \mathbf{88.6}\%$$

9A Describe Gas Laws and use to Calculate

Boyle's Law

This law says: The pressure and the volume of a gas are inversely proportional to each other, when temperature and moles of gas are kept constant. In other words, if you keep the temperature and moles of gas constant, and you increase the pressure on the gas, the volume will decrease, and if you lessen the pressure on a gas, the volume will increase.

$P_1 V_1 = P_2 V_2$
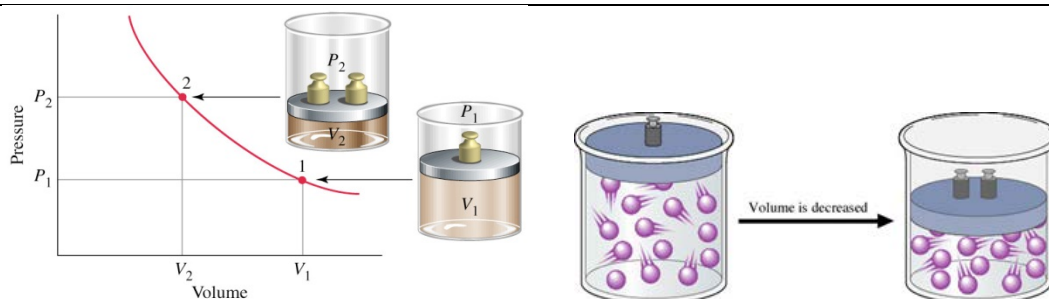
Charles' Law

This law says: The temperature and the volume of a gas are directly proportional, assuming pressure and moles of gas are kept constant. Another way to say this is: if the temperature goes up the volume will go up when pressure and number of moles are constant. The opposite also happens; if temperature is lowered, volume will be less. TEMPERATURE IN KELVIN ONLY

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$
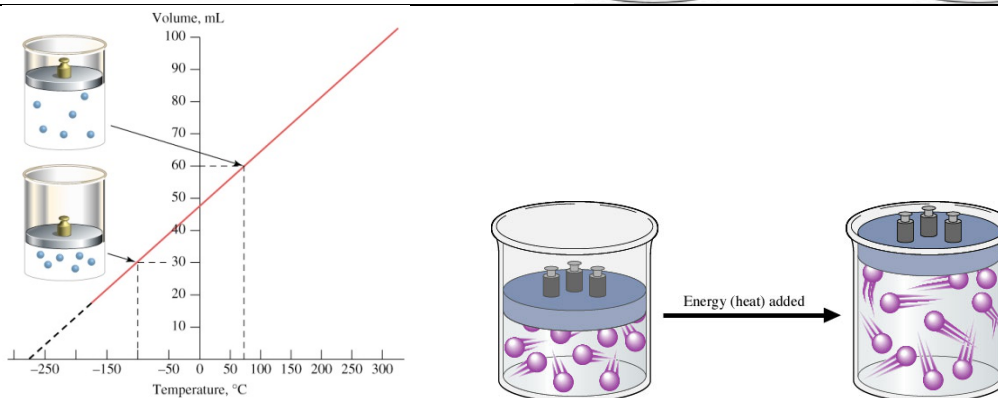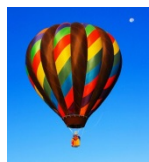
Avogadro's Law

Equal moles of gases at the same temperature and pressure will occupy equal volumes. Volume and moles are directly proportional if temperature and pressure are kept constant. Simply, more moles results in a larger volume and fewer moles results in a smaller volume if T and P are not changing.
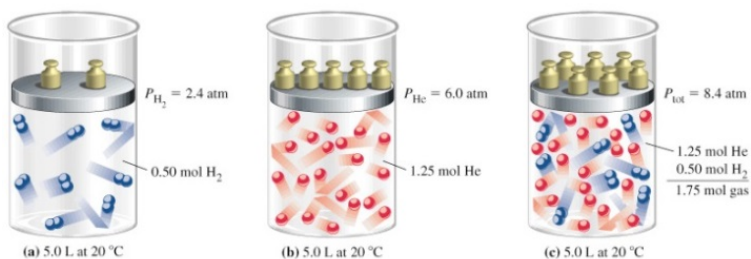
STP = $0^\circ C$, and at 1 atm. At STP, one mole of gas, any gas, will occupy exactly 22.4 liters.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

	He	N₂	CH₄
Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0°C	0°C	0°C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Dalton's Law of Partial Pressures total pressure = sum of all partial pressures

•partial pressure: pressure exerted by one gaseous component in a mixture

•Therefore, the pressure of a mixture of gases is equal to the sum of the partial pressures of the individual gases.

$P_{H_2} = 2.4$ atm   0.50 mol H_2 (a) 5.0 L at 20 °C

$P_{He} = 6.0$ atm   1.25 mol He   (b) 5.0 L at 20 °C

$P_{tot} = 8.4$ atm   1.25 mol He   0.50 mol H_2 1.75 mol gas   (c) 5.0 L at 20 °C

Combine Gas Law ; $\dfrac{P_2V_1}{T_1}=\dfrac{P_2V_2}{T_2}$ or $\dfrac{P_2V_1}{T_1n}=\dfrac{P_2V_2}{T_2n}$ USE WHEN CONDITIONS CHANGE

Ideal Gas Law

PV =nRT USE WHEN CONDITIONS DO NOT CHANGE. All Units must be the same units that are in R

P - pressure (in atm).

V - volume (in liters). Type equation here.

n - quantity of gas (in moles).

R - the gas constant .0821 atm L / mol K.

T - temperature, IN KELVINS ONLY ALWAYS

Example Problems:

1. A small 2.00 L fire extinguisher has an internal pressure of 506.6 kPa at 25°C. What volume of methyl bromide, the fire extinguisher's main ingredient, is needed to fill an empty fire extinguisher at standard pressure if the temperature remains constant?

This problem is solved by inserting values into (change in conditions)

$$\dfrac{P_1V_1}{T_1n_1} = \dfrac{P_2V_2}{T_2n_2} \text{ OR } P_1V_1 = P_1V_1 \text{ PLUG IN } (506.6 \text{ Kpa})(2.00\text{L}) = (101.3 \text{ Kpa})(V_2)$$

$$V_2 = \dfrac{506.6 \text{ Kpa})(2.00\text{L})}{101.3 \text{ Kpa}} = \mathbf{10.0L}$$

2. If a gas has a volume of 3.00 L, a pressure of 740.0 Torr and a temperature of 310.0 K, what will the temperature of the gas be in $^\circ$C if the pressure is increased to 1.50 atm while keeping the volume constant?

$V_1 = 3.00\text{L}$ $V_2 = 3.00\text{L}$ constant so V can be deleted from equation

$P_1 = 740 \text{ Torr}$ $P_2 = 1.50 \text{ atm} = \dfrac{1.50 \text{ atm}}{1} \times \dfrac{760 \text{ Torr}}{1 \text{ atm}} = 1.14 \times 10^3 \text{ Torr}$

$\qquad\qquad\qquad T_1 = 310K$ $T_2 = ?\ in\ C^o$

Change in conditions

$$\dfrac{P_1V_1}{T_1n_1} = \dfrac{P_2V_2}{T_2n_2} = \dfrac{P_1}{T_1} = \dfrac{P_2}{T_2} = P_1T_2 = P_2T_1 = T_2 = \dfrac{P_2T_1}{P_1}$$

$$T_2 = \dfrac{(1.14 \times 10^3 \text{ Torr})(310.\text{K})}{740. \text{ Torr}} = 478 \text{ K}$$

$\qquad\qquad = $ and we know: K= C^o + 273

$\qquad\qquad$ 478 K= C^o + 273   C^o = 478-273 = **205 C^o**

3. A sample of 6.9 moles of gas is placed in a container with a volume of 30.4 L at 62°C. What is the pressure of exerted by the gas?

$PV{=}nRT$ (no change in conditions) All units must match the units in the R you use. I am using .0821 atm L/mol   K= C^o + 273= 62°C + 273 =335 K
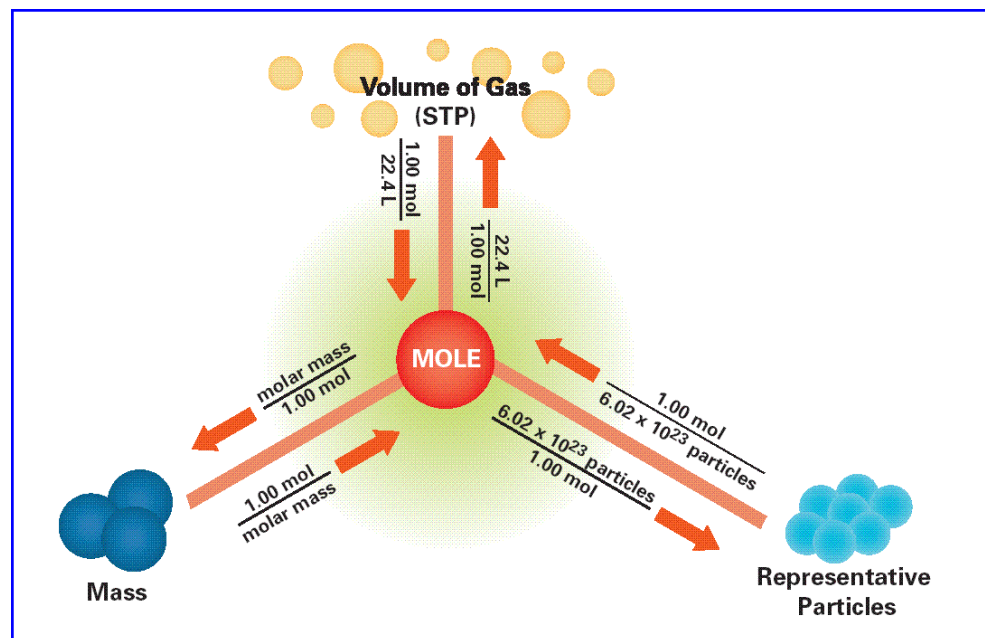
$$PV = nRT \quad P = \dfrac{nRT}{V} \quad \text{plug in} \quad P = \dfrac{(6.9 \text{ mol})(.0821\ \frac{\text{atm L}}{\text{mol K}})(335 \text{ K})}{30.4 \text{ L}} = \mathbf{6.2 \text{ atm}}$$

4. Three gases A, B, and C are contained in a tank. What is the pressure in atmospheres in the tank if the P_A is 3.5 atm., the P_B is .724 atm, and the P_C is .450 atm?

$P_T{=}P_1{+}P_2{+}P_3{+}...$ = P_T= 3.5 atm + .724 atm + .450 atm = **4.67 atm**

9B Gas Stoichiometry

Volume of Gas (STP)

$\dfrac{1.00\ mol}{22.4\ L}$

$\dfrac{22.4\ L}{1.00\ mol}$

MOLE

$\dfrac{molar\ mass}{1.00\ mol}$

$\dfrac{1.00\ mol}{molar\ mass}$

$\dfrac{6.02 \times 10^{23}\ particles}{1.00\ mol}$

$\dfrac{1.00\ mol}{6.02 \times 10^{23}\ particles}$

Mass

Representative Particles

Nitrogen monoxide and oxygen gas combine to form the brown gas nitrogen dioxide, which contributes to photochemical smog. What volume (in L) of nitrogen dioxide gas is produced when 34 L of oxygen gas react with an excess of nitrogen monoxide? Assume conditions are at STP.

$$2NO + O_2 \rightarrow 2NO_2$$
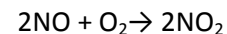
$$34L\ O_2 = \underline{\hspace{2cm}}L\ NO_2$$

What do we know?

$$1\ mol = 22.4\ L$$

$$1\ mol\ O_2 = 2\ mol\ NO_2$$

$$\frac{34\ \cancel{L\ O_2}}{1} \times \frac{1\ \cancel{mol\ O_2}}{22.4\ \cancel{L\ O_2}} \times \frac{2\ \cancel{mol\ N_2}}{1\ \cancel{mol\ O_2}} \times \frac{22.4\ L\ NO_2}{1\ \cancel{mol\ N_2}} = \boldsymbol{68\ L\ NO_2}$$

9C Kinetic Molecular Theory

THE KINETIC MOLECULAR THEORY is used to explain the behavior of gases and is based upon the following postulates:

1. Gases are composed of a many particles that behave like hard spherical objects in a state of constant, random motion.
2. These particles move in a straight line until they collide with another particle or the walls of the container.
3. These particles are much smaller than the distance between particles, therefore the volume of a gas is mostly empty space and the volume of the gas molecule themselves is negligible.
4. There is no force of attraction between gas particles or between the particles and the walls of the container.
5. Collisions between gas particles or collisions with the walls of the container are elastic. That is, none of the energy of the gas particle is lost in a collision.
6. The average kinetic energy of a collection of gas particles is dependent only upon the temperature of the gas.

In short, the key parts are:

1. constant random motion
2. straight line motion
3. negligible volume
4. no forces of attraction
5. elastic collisions
6. KE α T

gas molecule

container
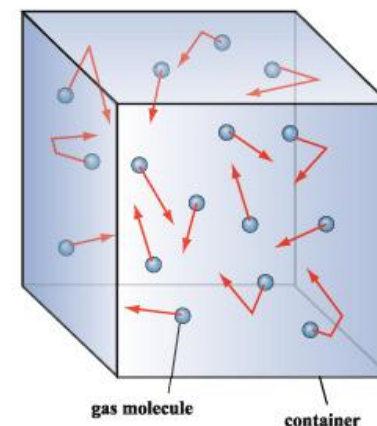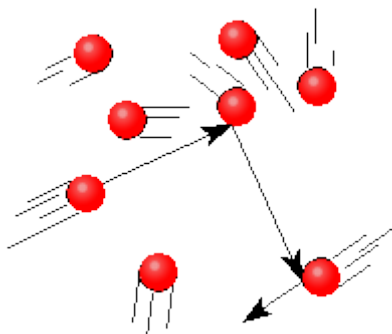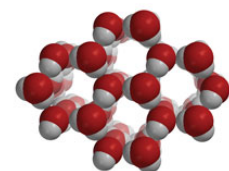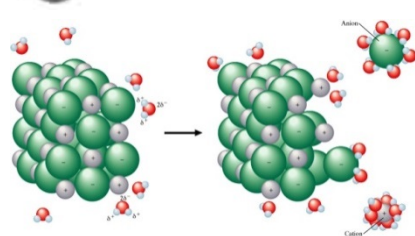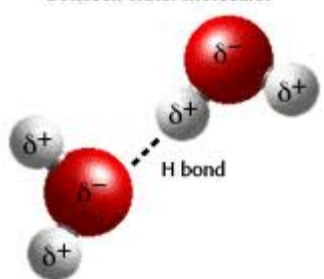
10A Water

1. Water molecules stick to each other—a force called cohesion. They also stick to other things—a force called adhesion. - Strong Surface Tension- Charged ends of water molecules are attracted to other substances with charges.

2. Water molecules pull on other water molecules. This pulling, called capillary action, allows liquid water to move against gravity. Capillary action is very important in nature. It makes it possible for water to travel from the roots of plants to the tips of their leaves—even to the tops of towering trees!

3. The stickiness of the hydrogen bonds also causes liquid water to form a kind of "skin"—called surface tension—where it meets the air

2. High Heat Capacity - Water's ability to absorb heat helps protect Earth from wild temperature swings from night to day and summer to winter. It keeps ocean temperatures fairly constant, so creatures that live there have a relatively stable environment. Water's ability to soak up heat as it changes from one form to another helps keep you cool too. When your sweat evaporates, it draws heat from your body.

3. Good solvent- Scientists call water the universal solvent. Rivers and streams carry nitrogen and phosphorus, which plants need to grow. They also carry pollutants, which can harm living things. Water carries nutrients from plants' roots to their leaves. It carries food through an animal's bloodstream to nourish its cells.

4. Water's ability to exist in all three forms of matter at temperatures common on Earth helps it move from one place to another.

5. Ice float on water. Unlike most substances the solid form of water is less dense than the liquid. When water freezes the molecules move part. If ice didn't float, then in winter the water at the top of the lake, near the cold air, would freeze and sink to the bottom.

10B. Example Developing Solubility Rules

1. After three drops of 1.0 M silver chloride is added to three drops of 1.0 M NaCl a white precipitate is observed.

2. After three drops of 1.0 M magnesium nitrate is added to three drops of 1.0 M NaCl no precipitate is observed.

Reaction 1 $AgNO_3$ + NaCl → AgCl + $NaNO_3$ The precipitate has to be either $NaNO_3$ or AgCl

Reaction 2 $Mg(NO_3)_2$ + 2NaCl → $MgCl_2$ +2 $NaNO_3$ Both products are soluble.

We now know $NaNO_3$ is soluble so AgCl must be the insoluble precipitate formed in reaction 1.

Example Using Solubility Rules

By using your solubility rules, help us figure out which numbered solution matches up with the following chemicals: $AgNO_3$, NaCl, KNO_3, NaOH

You will be given a labeled solution of $CuSO_4$ to help you with this task.

Data Table

	1	2	3	4
$CuSO_4$	No reaction	No reaction	Ppt formed	No reaction
1	X	No reaction	Ppt formed	Ppt formed
2	No reaction	X	No reaction	No reaction
3	Ppt formed	No reaction	X	Ppt formed
4	Ppt formed	No reaction	No reaction	X

Sample Conclusion

Here is a brief description of how the students should come up with the identity of the unknowns:

Solution 3 is NaOH because it is the only one of the four that will form a precipitate with $CuSO_4$.

Solution 2 is KNO_3 because it does not form a precipitate with anything.

Solution 1 is $AgNO_3$ because it forms a precipitate with Solution 3, which we have already determined is NaOH.

Solution 4 is NaCl because it forms a precipitate with $AgNO_3$, which is Solution 1.

10C Molarity

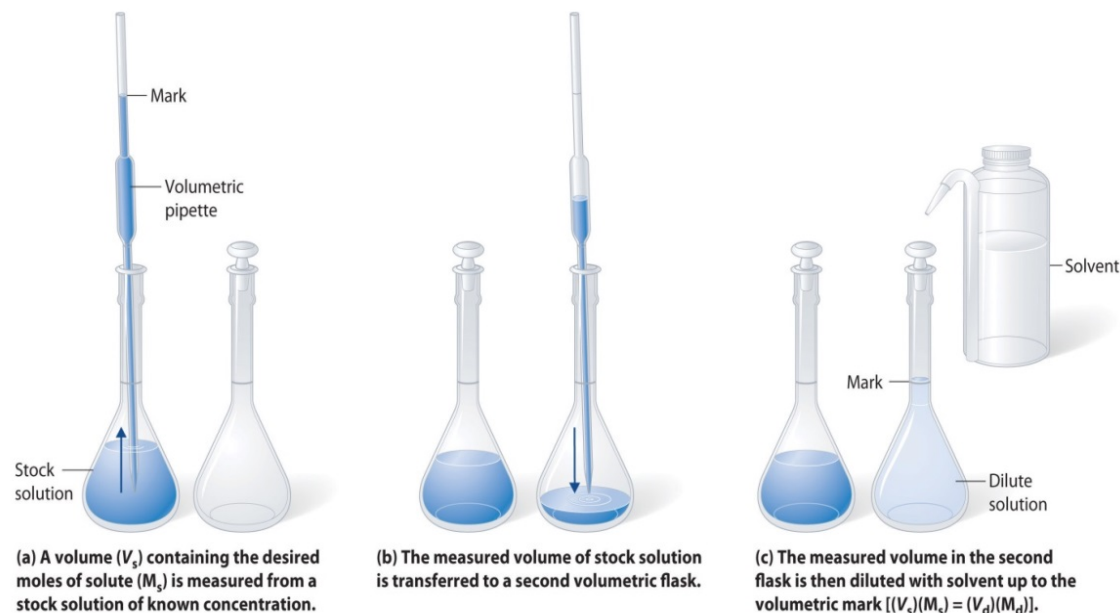The definition of molarity (M) is $M = \frac{mole\ of\ solute}{L\ of\ solution}$

Example:

Calculate the molarity of a solution containing 8.61 g potassium hydroxide, dissolved in enough water to form 750 mL of solution. You to use M formula 5.61 g of KOH will need to be converted to mol and 750 mL will need to be converted to L.

$$\frac{86.1\ \cancel{g\ KOH}}{1} \times \frac{1\ mol\ KOH}{56.105\ \cancel{g\ KOH}} = 1.53\ mol\ KOH$$

$$\frac{750\ ml\ of\ solution}{1} \times \frac{1\ L}{1000\ \cancel{mL}} = .750\ L\ of\ solution$$

$$M = \frac{mole\ of\ solute}{L\ of\ solution} = \frac{1.53\ mol\ KOH}{.750\ L\ solution}$$

$$= 0.014\ M\ or\ \frac{mol}{L}$$

Cobalt(II) chloride dihydrate
$(CoCl_2 \cdot 2H_2O)$

Ethanol

500 mL

Mark

500 mL

500 mL

500 mL

100.0 g

(a) An amount of solute is weighed out on an analytical balance and then

(b) A portion of the solvent is added to the volumetric flask.

(c) The mixture is swirled until *all* of the solute is dissolved.

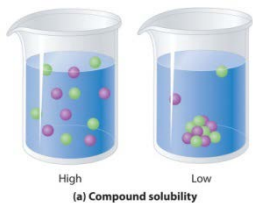**(d)** Additional solvent is added up to the mark on the volumetric flask.

10D Dilutions $$M_1V_1 = M_2V_2$$

Concentrated H_2SO_4 is 18.0 M. What volume is needed to make 2.00 L of 1.00 M solution?

$$M_1V_1 = M_2V_2$$

$$18.0\ M \times V_1 = 1.00\ M \times 2.00L$$

$$V_1 = \frac{1.00\ \cancel{M} \times 2.00L}{18.0\ \cancel{M}} = .111\ L\ or\ 111\ mL$$

Mark

Volumetric pipette

Solvent

Mark

Stock solution

Dilute solution

(a) A volume (V_s) containing the desired moles of solute (M_s) is measured from a stock solution of known concentration.

(b) The measured volume of stock solution is transferred to a second volumetric flask.

(c) The measured volume in the second flask is then diluted with solvent up to the volumetric mark [(V_s)(M_s) = (V_d)(M_d)].

10E Types of Soltions

Solubility is defined as the maximum amount of solute dissolved by a given amount of solvent at a specified temperature. The solubility of the given substances in a given solute is temperature-dependent.

Unsaturated: If the solution contains less quantity of solute than what can be dissolved, at a specified temperature it is unsaturated. When more solute is added into solution the solute dissolves.

An unsaturated solution is one in which more of the solute could dissolve at the same temperature.

Saturated: A saturated solution is one in which no more of the solute will dissolve at a specific temperature. Solution containing the maximum amount of solute at a specified temperature is saturated. When more solute is added into the solution the solute will no longer dissolve.

Supersaturated: A supersaturated solution is when a solution which contains more solute than would normally dissolve at a certain temperature. The solution becomes supersaturated by elevating the temperature and the solution contains the maximum amount of solute at an elevated temperature is carefully cooled to a lower temperature and is now supersaturated. When more solute is added into the solution, crystals will form.

10F Solubility is the maximum amount of a substance that can be dissolved in a given amount of solvent. Solubility depends on temperature and pressure, as well as on the type of solvent being used. Solubility is not affected by an increase in the surface area of the solute or by agitation of the solution although the rate of solution formation (dissolution) is affected.
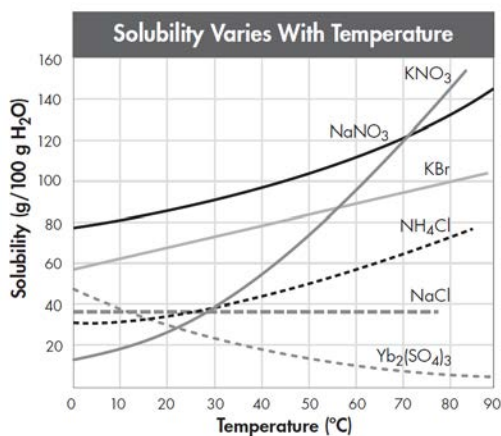
Factors Affecting Solubility:

Temperature affects solubility of solids differently than the solubility of gases.

Solid solubility: ↑ temperature often (not always) ↑ solubility
Solid pressure has no effect on solubility

Gas solubility: ↑ temperature ↓ solubility
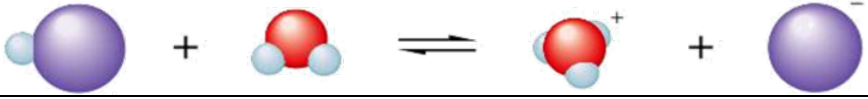**Gas** solubility: ↑ pressure ↑ solubility

Factors Affecting Rate of Solution Formation (dissolution) of Solids:
↑ temperature ↑rate of dissolution
↑ agitation (stirring or shaking) ↑rate of dissolution
↑surface area (grinding or crushing a solid) ↑rate of dissolution

10G Acid Base Theories

ARRHENIUS DEFINITION	BRØNSTED-LOWRY DEFINITION
An acid is a substance that contains a hydrogen and ionizes to produce hydrogen ions in aqueous solutions. **Acid-**produces a proton (H^+) in water $HCl \xrightarrow{H_2O} H^+ + F^-$	Brønsted-Lowry acid is defined as a hydrogen-ion donor **Acid**-donates (gives up)a proton (H^+) $HF + H_2O \leftrightarrow H_3O+ + F^-$
A base is a substance that contains a hydroxide and dissociates to produce hydroxide ions in water (OH^-) **Base**-produces a hydroxide ion in water (OH^-) $NaOH \xrightarrow{H_2O} Na^+ + Cl^-$	Brønsted-Lowry base is defined as a hydrogen-ion acceptor. **Base**-accepts (bonds to H^+) a proton (H^+) $NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$
require acid-base reactions to occur in aqueous solutions	
limited to substances with those "parts"; ammonia is a MAJOR exception!	Brønsted-Lowry definition includes all Arrhenius bases that produce hydroxide ions that can accept a proton. But the Brønsted-Lowry definition also covers other bases, such as ammonia, NH_3, that do not contain hydroxide ions.

10H Recognizing Neutralization, Precipitation and Oxidation and Reduction Reactions

Acid-base reactions are reactions that occur between an acid and a base in which a proton is transferred. The product includes a salt and, often, water.	Reactant an acid (formula starts with an H) and a base (formula end in OH or is NH_3). If the base ends in OH the products are a salt (first element a metal) and water. Salt has the + ion of the metal and - ion of the base. If the base is NH_3 product will only be a salt that has a formula that starts with NH_4 and the last ion in the formula will be the anion in the acid.
3.Precipitation reactions are reactions that occur when two aqueous solutions react and produce a solid precipitate.	Use solubility chart: If one product is insoluble the reaction is a precipitation reaction.
An **oxidation-reduction (redox) reaction,** also called a *redox* reaction, is any reaction that involves the exchange of an electron or oxygen. In general, when the oxidation number of a certain atom is different in one of the reactants than it is in one of the products, it means that at least one electron(s) were exchanged.	If the reaction if not an acid base or precipitation reaction it will most likely be a redox reaction. Assign oxidation numbers to all the elements in the equation if an elements oxidation number changes the reaction is a redox reaction.

10H
Oxidation Number Rules
1. The oxidation number for an atom in its elemental form is always zero. The oxidation number of a monoatomic ion = charge of the monatomic ion.
2. The oxidation number of all Group 1A metals = +1 (unless elemental) and The oxidation number of all Group 2A metals = +2 (unless elemental).
3. In ionic binary compounds the oxidation number is the same as the oxidation number used to write the formula of the compound
4. Hydrogen (H) has two possible oxidation numbers:
+1 when bonded to a nonmetal
-1 when bonded to a metal
5. Oxygen (O) oxidation numbers is -2:
6. The oxidation number of fluorine (F) is always -1.
7. The sum of the oxidation numbers of all atoms (or ions) in a neutral compound = 0.
8. The sum of the oxidation numbers of all atoms in a polyatomic ion = charge on the polyatomic ion.

10I Calculating pH

$K_w = [H_3O^+][OH^-]$ $K_w = 1 \times 10^{-14}$ at 25°C
$pH = -\log [H_3O^+]$
$pOH = -\log [OH^-]$
$1 \times 10^{-14} = [H_3O^+][OH^-]$
$14 = pH + pOH$
$[H_3O^+] = 10^{-(pH)}$
$[OH^-] = 10^{-(pOH)}$

**Example 1**: calculate the pH of a solution with a hydrogen ion concentration of 6.3×10^{-4} M

$$pH = -\log (6.3 \times 10^{-4})$$
$$= -(-3.2) = \textbf{3.2}$$

Example 2 if a solution has a hydroxide ion concentration of 1.3×10^{-10}, then you can calculate the pOH and find the pH.

$$pOH = -\log (1.3 \times 10^{-10}) = 9.9$$
$$\text{Therefore, pH} = 14 - 9.9 = \textbf{4.1}$$

10J The strength of acids and bases depends on the degree to which they dissociate. The greater the degree of dissociation, the greater their strength.
In general, strong acids and strong bases dissociate, or ionize, completely in aqueous solution. In other words, almost 100 percent of a strong acid or strong base interacts with water and ionizes. A strong dissociation is described using one arrow pointing to the right, as shown.
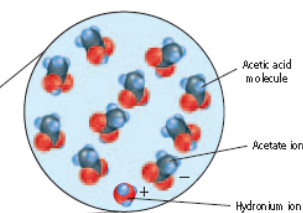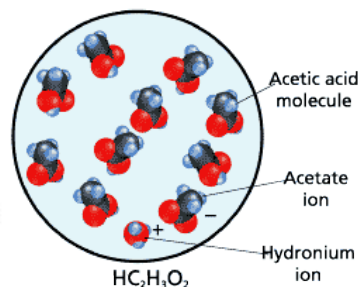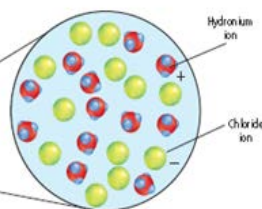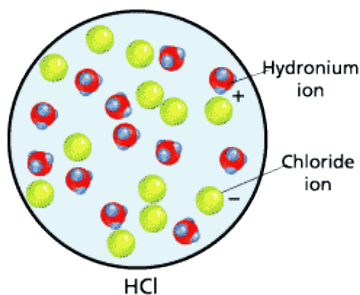
Strong Acid →   All of HCL becomes H_3O^+ and Cl^- Weak Acid → very small amount of $HC_2H_3O_2$ becomes H_3O^+ and $C_2H_3O_2^-$

11A Concept of Energy
1. The Usual Definition of Energy: the ability to do work and to transfer heat.
work is moving an object against an opposing force
work = distance × opposing force $w = d \times f$
SI unit of work or energy: the joule (J)
2. Two Basic Forms of Energy: potential energy: energy of position   examples: boulder on a ledge, chemical bonds
kinetic energy: energy of motion
examples: pool balls, molecules
Forms of Energy:
Energy is found in different forms including light, heat, chemical, and motion. There are many forms of energy, but they can all be put into two categories: potential and kinetic.

11B Conservation of Energy

Whether a reaction absorbs energy or releases energy, there is no overall change in the amount of energy. Energy cannot be created or destroyed. This is the **law of conservation of energy**. Energy can change form — for example, from electricity to light — but the same amount of energy always remains.

If energy cannot be destroyed, what happens to the energy that is absorbed in an **endothermic reaction**? The energy is stored in the chemical bonds of the products. This form of energy is called chemical energy. In an endothermic reaction, the products have more stored chemical energy than the reactants. In an **exothermic reaction**, the opposite is true. The products have less stored chemical energy than the reactants. The excess energy of the reactants is released to the surroundings when the reaction occurs. The following graphs compare the energy changes in endothermic and exothermic reactions.
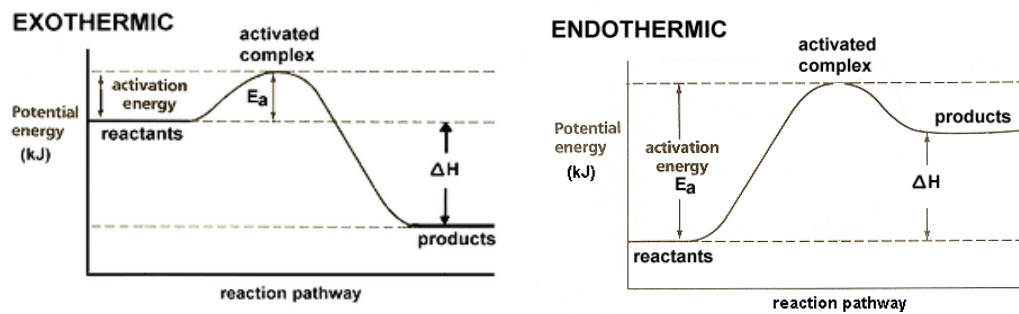
11C Calculate Enthalpy of Reaction ΔH$_{rxn}$ from Bond Energies If not given you will need to write the Lewis structures to use bond energy data to determine ΔH$_{rxn}$.
The bond enthalpy is the energy needed to break one mole of bonds in gaseous molecules under standard conditions (298 K and 1 atm).
The bond enthalpy is the energy needed to break one mole of bonds in gaseous molecules under standard conditions (298 K and 1 atm).

Heat Flow	Sign of q	Type of Reaction	Products	Surroundings
from surroundings to system	+ (positive)	endothermic	gain heat	lose heat (get colder)
from system to surroundings	– (negative)	exothermic	lose heat	gain heat (get hotter)

Example

The complete combustion of propane can be represented by the following equation:

$CH_3CH_2CH_3(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4 H_2O(g)$
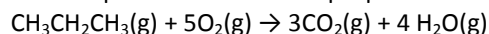
or we should redraw the equation for the reaction to represent the bonds present:
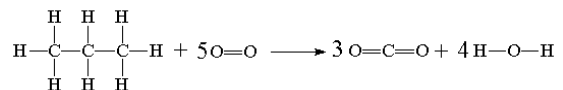
Total endothermic change for bond breaking:

 8 x C-H   2 x C-C   5 x O=O

(8 x +413) + 2(+347) + 5(+498)

Total exothermic change for bond forming:

 6 x C=O   8 x O-H

(6 x -805) + (8 x -464)
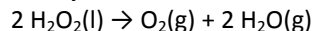
ΔH= [(8x413)+(2x347)+(5x498)] - [(6x-805)+(8x-464)]=-**2054 kJ/mol**

- ΔH so the reaction **is exothermic** releasing energy to the surrounding's

Calculate Enthalpy of Reaction ΔH$_{rxn}$ from Standard Enthalpies of Formation
- Molar Heat of Formation (ΔH$_f$): is the amount of heat required / given off to make 1 mole of a compound from its elemental components. Enthalpy of Formation (also called Heat of Formation)
- Molar Heat of Formation of **ALL ELEMENTS** is 0 kJ.
- The state of the compound affects the magnitude of ΔH$_f$. (H$_2$O (g) has ΔH$_f$ = −241.8 kJ/mol; H$_2$O (l) has ΔH$_f$ = −285.8 kJ/mol)
- Standard Molar Heat of Formation (ΔH°$_f$) is the heat of formation under standard conditions (1 atm and 25°C).
- Enthalpies of formation for many chemical substances can be found in tables.

$$\Delta H = \sum n\Delta H^{o}_{f\ product} - \sum n\Delta H^{o}_{f\ reactant}$$

Example: Calculate the enthalpy of reaction for decomposition of hydrogen peroxide into oxygen and water and determine if the reaction is exothermic or endothermic.
2 H$_2$O$_2$(l) → O$_2$(g) + 2 H$_2$O(g)
ΔH$_f$ of reactants:
ΔH$_f$ H$_2$O$_2$(l) =-187.8 kJ/mol
ΔH$_f$ of products:
ΔH$_f$ O$_2$(g) = 0.0 kJ/mol
ΔH$_f$ H$_2$O(g) =-285.8 kJ/mol
Using the balanced equation:
ΔH = [0.0 + 2 mol (-285.8 kJ/mol)] - [2 mol (-187.8 kJ/mol)] = -**196 kJ/mol**   Because ΔH = -196 kJ/mol is negative, the reaction is **exothermic**

11D Calorimetry
Units of Energy:
calorie--amount of heat needed to raise the temp. of 1.00 gram of water 1.00 °C
kilocalorie—1cal ==1000 Kcal the food calorie with a capital C.
joule--SI unit of energy; 1 cal = 4.184 J
Specific heat,(s) also known as **specific heat capacity (Cp) (sometimes referred to as just heat capacity),** is defined as the amount of heat necessary to raise the temperature of 1 g of a substance by one degree. joules per gram per degree, J/g°C or J/g K. The change in temperature Δt =t$_{final}$-t$_{initial}$ is the same whether you use °C or K.
Q=(mass) (Cp) (Δt)
Specific heat of water (liquid state) = 4.184 J/g°C (or 1.00 cal/g°C)
Heat(enthalpy) of fusion is the amount of heat required to change a specified amount of a solid to a liquid or released to change a given amount of a liquid to a solid at the melting or freezing point of the substance.
Heat(enthalpy) of vaporization is the amount of heat to change a given amount of substance form an liquid to a gas or released to change a given amount of a gas to a liquid at the substance boiling point.
Calorimetry: is a method of measuring heat flow between systems and surroundings.
The device used to measure heat flow between systems and surroundings is called a calorimeter
   Q=(mass) (Cp) (Δt) change in temperature

Thermometer
Glass stirrer
Cork stopper
Two Styrofoam ®
cups nested
together containing
reactants in solution

Example 1: How much heat is required to raise 80.0 g of aluminum from 12°C to 46°? The specific heat capacity of aluminum is 0.90kJ/kg°C.
Q is heat so we are solving for Q. we can just plug into: Q = (m) (C) (Δt), If we use 0.90kJ/kg°C we have to use kg
(80 g)/1×(1 kg)/1000g=0.080 kg.
Q= (0.080kg)(0.90kJ/kg°C)(46°C-12°C) = **2.488 kJ**
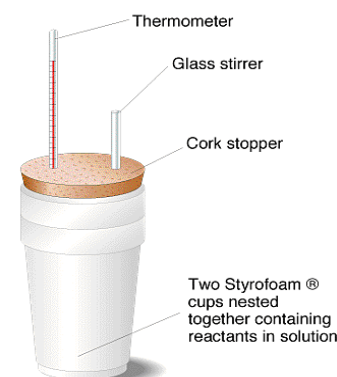
Example 2: How much water at 25°C must be added to 15 kg of water at 10 °C so that the mixture becomes 20°C? Qp = (m) (C) (Δt)

Q hotter water = Q of cooler water LCE: heat will be transferred from the hotter water to cooler water until they are both at the same temperature.

Absolute change in temperature for the hotter water is $25^{\circ}C - 20^{\circ}C = 5^{\circ}C$

Absolute change in temperature for the cooler water is = $20^{\circ}C - 10^{\circ}C = 10^{\circ}C$

Q hotter = Q cooler

(m) (cp) (Δt)= (m) (C) (Δt)

(m) (4.184 J/g°C) (5°C) = (15 kg) (4.184 J/g°C) (10°C)

$$m = \frac{(15\ kg)\ (4.184\ J/g°C)\ (10°C)}{(4.184\ J/g°C)\ (5°C)}$$

=30 kg

Example 3: We are going to determine the specific heat of a metal using experimental data. A piece of hot metal is placed in cooler water. The heat from the hotter metal will move to the cooler water until both the metal and the water have the final temperature. In this experiment, we used a "coffee cup" calorimeter and gathered the following data:

Mass of empty cup	2.31 g
Mass of cup + water	80.89 g
Mass of cup + water + metal	780.89 g
Initial temperature of water	17.0 °C
Initial temperature of metal	52.0 °C
Final temperature of system	27.0 °C

The key thermochemistry equation for solving this problem is:

Q metal = Q water

(mass) (Δt) (C) = (mass) (Δt) (C)

Then, by substitution, we have (metal values on the left, water values on the right):

We need to use the data table to get what we need to substitute into the above equation.

     mass of water: 180.98 g cup + water - 2.31g cup = 178.58 g water

     mass of metal: 780.89 g cup + water + metal - 180.89 g cup + water = 600.0 g metal
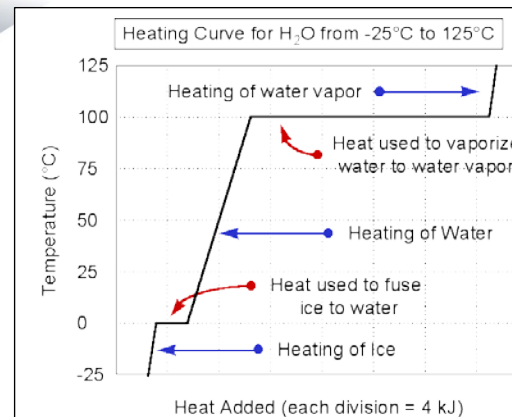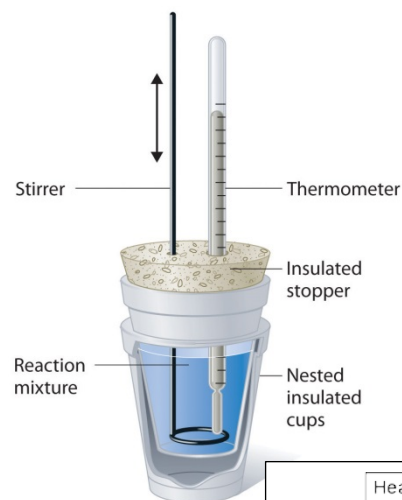
     change in water temperature: 27.0 °C - 17.0 °C = 10.0 °C

     change in metal temperature: 52.0 °C - 17.0 °C = 25.0 °C

Putting the numbers in place gives us:

(600.0 g) (25.0 °C) (C) = (178.58 g) (10.0 °C) (4.184 J/g°C)

$$C = \frac{(178.58\ g)\ (10.0\ °C)\ (4.184\ J/g°C)}{(600.0\ g)\ (25.0\ °C)\ (C)} = 0.498\ \frac{J}{g°C}$$

Stirrer — Thermometer

Insulated stopper

Reaction mixture — Nested insulated cups

Heating Curve for H₂O from -25°C to 125°C

Heating of water vapor

Heat used to vaporize water to water vapor

Heating of Water

Heat used to fuse ice to water

Heating of Ice

Heat Added (each division = 4 kJ)

11E ΔH$_{rxn}$ ΔH$_{rxn}$ values are derived from experimental data. Experiments that measure heat exchange are referred to collectively as calorimetry experiments, and the measurement device is called a calorimeter. There are two important types of calorimetry experiments: constant-pressure calorimetry (also known as **coffee cup calorimetry**) and constant-volume calorimetry (also known as **bomb calorimetry**).

Measuring Heat of Reaction Experimentally

i. A known quantity of reactant is placed in a well-insulated vessel (a Styrofoam cup)

ii. The initial temperature of this reactant is recorded, t$_i$.

iii. A known quantity of the second reactant is added, the vessel is sealed with a lid and the reaction mixture stirred.

iv. The final temperature of the reaction mixture is recorded, t$_f$.

v. The the heat released or absorbed, in joules, for the reaction is calculated:

Q=(mass) (Δt) (C) Δt = t$_f$ - t$_i$

vi. The enthalpy change in kJ per mole of a given reactant for the reaction is calculated:

$$\Delta H = \frac{kJ}{mole}$$

Common assumptions for reaction mixtures made up of aqueous solutions:

- density of aqueous solution assumed to be the same as for water, 1g/mL , 100mL of solution is said to have a mass of 100g
- additivity of volumes of reactants is assumed, 100mL of reactant a + 200mL of reactant b = 300mL of reaction mixture
- specific heat capacity of the reaction mixture assumed to be the same as water, specific heat capacity = 4.184 J/g^{o}C

Heat is not lost to, or absorbed by, the surroundings.

Example Heat (Enthalpy) of Solution

Ammonium chloride is very soluble in water. When 4.50 g NH$_4$Cl is dissolved in 53.00 g of water, the temperature of the solution decreases from 20.40 °C to 15.20 °C. Calculate the enthalpy of dissolution of NH$_4$Cl (in kJ/mol). (or enthalpy or heat of solution) You are asked to calculate the enthalpy change for the dissolution of ammonium chloride. You are given the masses of the solid and water, and the temperature change that occurs when the two are combined. First calculate the energy change for the surroundings (Q solution) in the coffee cup calorimeter.

Q solution = m(solution) × C solution × ΔT

C$_p$ solution = 4.18 J/g • °C (always assume same as water for coffee cup calorimetry)

m solution = 4.50 g + 53.00 g = 57.50 g

ΔT = T final – T initial = 15.20 °C - 20.40 °C = -5.20 °C

Q solution = m(solution) × C solution × ΔT

Q = (57.50 g̶)(4.18 J/g̶ × °C̶)(-5.20 °C̶) = **-1250 J** The surroundings lost energy is negative so the system gained energy and the reaction is endothermic and Δ*H* is positive

Q dissolution = 1250 J

△H has the units KJ /mole so we will convert 1250 J to kJ and 4.50 g NH$_4$Cl to moles to determine △H

$$\frac{1250\ \cancel{J}}{1} \times \frac{1 kJ}{1000\ \cancel{J}} = 1.250\ kJ \qquad \frac{4.50\ \cancel{g\ NH_4Cl}}{1} \times \frac{1\ mole\ NH_4Cl}{53.492\ \cancel{g\ NH_4Cl}} = \mathbf{0.0841\ mole\ NH_4Cl}$$

$$\triangle H = \frac{1.250\ kJ}{0.0841\ mol\ NH_4Cl} = \textbf{14.86 kJ/mol}$$

12A Radiation

	Alpha	Beta	Gamma
Symbol	α	β	γ
Composed of...	2 protons, 2 neutrons	1 electron	1 photon
A.K.A.	Helium nucleus	Fast moving (high energy) electron	Electromagnetic (EM) Radiation
Chemical Symbol	^4_2He	$^0_{-1}e$	$^0_0\gamma$
Charge (e)	+ 2	- 1	0
Mass (kg)	6.64×10^{-27}	9.11×10^{-31}	0
Ionising Power	Large	Small	Tiny
Absorbed by...	Paper, Skin	Al (~ 5mm) Pb (2-3mm)	Concrete (> 1m) Pb (several cm)
Range in air	~ 5cm	~ 1m	∞
Magnetic Field	deflected	deflected (opposite to α)	unaffected
Electric Field	deflected	deflected (opposite to α)	unaffected

Types of Radiation you need to know:

A= mass number which is the sum of the neutrons and protons

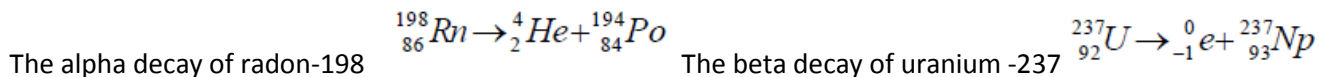| Radiation	Type of Radiation	Mass (AMU)	Charge	Shielding material
Alpha	Particle	4	+2	Paper, skin, clothes
Beta	Particle	1/1836	±1	Plastic, glass, light metals
Gamma	Electromagnetic Wave	0	0	Dense metal, concrete, Earth

Z= atomic number which is the number of protons
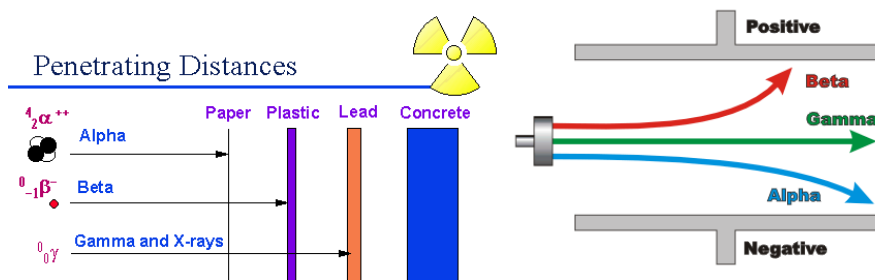
PROCESS	CHANGE IN A	CHANGE IN Z	CHANGE IN NEUTRON/PROTON RATIO
ALPHA EMISSION	-4	-2	INCREASE
BETA EMISSION	0	1	DECREASE
GAMMA	0	0	X

Examples of Particles

beta particle (electron)	$_{-1}^{0}e$
positron	$_{+1}^{0}e$
alpha particle	$_{2}^{4}He$
proton	$_{1}^{1}H$
neutron	$_{0}^{1}n$
gamma ray	$_{0}^{0}\gamma$

12B Nuclear Reactions

Writing nuclear reactions requires that the atomic number, Z, and the mass number, A, add up to the same value on each side. In earlier chapters, Z stood for the number of protons. The mass number, A, is the sum of protons and neutrons. The mass of a proton or a neutron is about 1 amu.

$$_{86}^{198}Rn \rightarrow _{2}^{4}He + _{84}^{194}Po$$

The alpha decay of radon-198

The beta decay of uranium -237 $\quad _{92}^{237}U \rightarrow _{-1}^{0}e + _{93}^{237}Np$

Nuclear Transformation = converting one element into another by force.

* bombarding with alpha particles:

14/7N + 4/2He   1/1H + 17/8O +

* bombarding with neutrons:

238/92U + 1/1n   ☐ 239/92U

One way to identify the missing component of a nuclear reaction is to use the requirement that the total atomic number, Z, mass number,A, (the sum of the protons and neutrons', and so also called the nucleon number) and total charge Q must balance between reactants and products.

Example Nuclear Transformation

$^{249}Cf + {}^{18}O \rightarrow ? + 4\,n.$ The first thing to do is to write each component of the reaction in the form $\frac{atomic\ number}{mass\ number}$ X^{charge}, $_{98}^{249}Cf^{98+} + _{8}^{18}O^{8+} \rightarrow _{a}^{b}X^{c} + 4\,_{0}^{1}n^{0}$, where for now the unknown component is labeled X.

Next, we consider in turn the balance of atomic number (number of protons), mass number (number of protons plus neutrons), and total charge.

- The atomic number balance means:

$98 + 8 \rightarrow Z + 4 + 0$ and so the number of protons in the unknown species is $Z = 98 + 8 - 0 = 106$. From the atomic number we can immediate identify that the unknown nucleus is that of transuranium element with Z = 106, known as seaborgium, Sg.

- The mass number balance means:

$249 + 18 \rightarrow A + 4 \mu 1$ and so the mass number of the unknown species is A = 249 + 18 - 4 = 263. This means the Sg nucleus contains 106 protons and 263 - 106 = 157 neutrons.

NUCLEAR FISSION	NUCLEAR FUSION
A heavy nucleus breaks up to form two lighter nuclei.	Two light nuclei combine to form a heavy nucleus.
It involves a chain reaction.	Chain reaction is not involved.
The heavy nucleus is bombarded with neutrons.	Light nuclei are heated to an extremely high temperature.
We have proper mechanisms to control fission reaction for generating electricity.	Proper mechanisms to control fusion reaction are yet to be developed.
Disposal of nuclear waste is a great environmental problem.	Disposal of nuclear waste is not involved.
Raw material is not easily available and is costly.	Raw material is comparatively cheap and easily available.
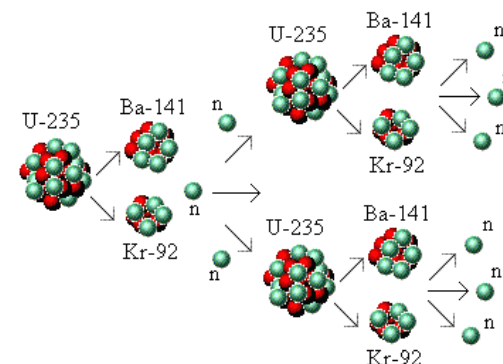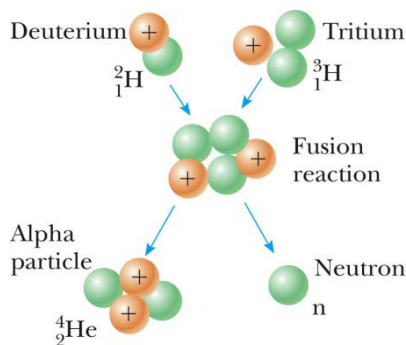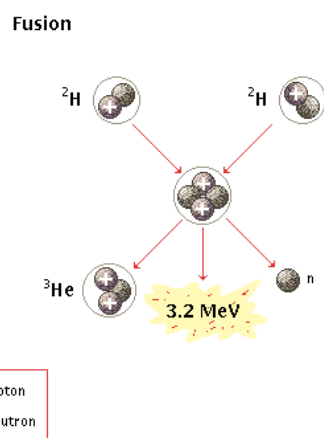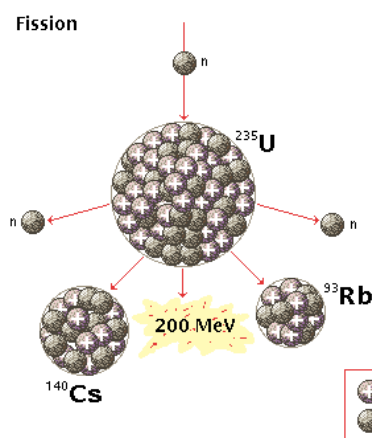
12C Fission vs. Fusion
There are two main types of nuclear reactions that can be discussed:
Fission: The process of causing a larger atom to split into multiple smaller atoms, releasing energy in the process.
- This is the system that we use in nuclear power plants.
- It is relatively easy to do, but also leaves us with lots of nuclear waste that must be stored safely for thousands of years.

Fusion: The process of causing smaller atoms to fuse together into a larger atom, in the process releasing energy.
- This is the process that drives our sun, and all other suns.
- We can do it under the right conditions in a lab, but we end up putting in more energy than we get out.
- The left over products of these reactions are safe, which is why a lot of research is going into developing fusion reactors.

Mnemonic	prefix	symbol	number	power of 10
King	kilo	k	1 000	10^3
Henry	hecto	h	100	10^2
died	deka	da	10	10^1
by	base unit	b	1	10^0
drinking	deci	d	0.1	10^{-1}
chocolate	centi	c	0.01	10^{-2}
milk	milli	m	0.001	10^{-3}

Expanded List of Metric Prefixes

prefix	symbol	number	power of 10
yotta	Y		10^{24}
zetta	Z		10^{21}
exa	E	1 000 000 000 000 000 000	10^{18}
peta	P	1 000 000 000 000 000	10^{15}
tera	T	1 000 000 000 000	10^{12}
giga	G	1 000 000 000	10^9
mega	M	1 000 000	10^6
kilo	k	1 000	10^3
hecto	h	100	10^2
deka	da	10	10^1
base unit	b	1	10^0
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000 001	10^{-6}
nano	n	0.000 000 001	10^{-9}
pico	p	0.000 000 000 001	10^{-12}
femto	f	0.000 000 000 000 001	10^{-15}
atto	a	0.000 000 000 000 000 001	10^{-18}
zepto	z	0.000 000 000 000 000 000 001	10^{-21}
yocto	y	0.000 000 000 000 000 000 000 001	10^{-24}

TABLE 1.4 SI Base Units

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s[a]
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

[a] The abbreviation sec is frequently used.