Chemistry Review 2007-2008

Synthesis – chemical combining of two or more smaller substances into a bigger one.

Endothermic – energy absorbed in a reaction

Exothermic – energy given off by a reaction

Precise – getting the same measurements over and over

Accurate – measurement that are almost exactly correct

Quantitative – numbers and units

Qualitative – description (fat, tall)

Density = mass/volume

Five pennies have a total mass of 14.5g, volume of 1.75 cm^3 find the density.

$D=M/V \quad \frac{14.5g}{1.75\ cm^3} = 8.2857142\ g/cm^3 = 8.26\ g/cm^3$ with three significant figures

K = C + 273

~Dimensional Analysis

How many seconds are in one year?

$1\ \text{year} \times \frac{365\ \text{days}}{1\ \text{year}} \times \frac{24\ \text{hrs}}{1\ \text{day}} \times \frac{60\ \text{min.}}{1\ \text{hr.}} \times \frac{60\ \text{sec.}}{1\ \text{min.}} = 31{,}536{,}000$ seconds

How many switches make 1 house ?

1 house = 8 rooms   1 room = 3 windows   1 window = 2 lights   3 lights = 4 switches

$1\ \text{house} \times \frac{8\ \text{rooms}}{1\ \text{house}} \times \frac{3\ \text{windows}}{1\ \text{room}} \times \frac{2\ \text{lights}}{1\ \text{window}} \times \frac{4\ \text{switches}}{3\ \text{lights}} = \frac{192}{3} = 64$ switches

~Scientific Notation

$2.100 \times 10^3 = 2.100$ $6.51 \times 10^{-5} = 0.0000651$

Addition $8.4 \times 10^7 + 1.1 \times 10^6 =$ $8.4 \times 10^7 + .11 \times 10^7 = 8.5 \times 10^7$

Multiplication $(5.0 \times 10^4)(3.0 \times 10^2) = 15 \times 10^6 = 1.5 \times 10^7$

Division $\frac{9.0 \times 10^{15}}{3.0 \times 10^3} = 3.0 \times 10^{12}$

Liquid mixtures – Solutions – Homogeneous

Solid to liquid is melting; liquid to solid is freezing
Liquid to gas is vaporization; gas to liquid is condensation
Solid to gas is sublimation; gas to solid is deposition

Homogeneous – same throughout

Heterogeneous – not the same

Element – can't be broken down, specific properties

Compound – 2 or more elements chemically combined, with specific properties different than atoms it's made up of

~

LAW OF CONSERVATION OF MATTER

Matter cannot be created or destroyed by a chemical reaction, or in a physical change.

solids – have a definite shape and volume

liquids – have indefinite shape with a definite volume

gases – have an indefinite shape and volume

Chemical Properties	Physical Properties
how atoms or compounds chemically combine or decompose, types of ions formed,	color, solubility, odor, hardness, density, melting point, freezing point, H_F, H_V, C, etc.

Indicators of a chemical reaction: TOPIC—B

T – temperature change
O – odor
P – precipitate
I – irreversibility
C – color change
B- bubbles

reactants – starting point chemicals

products – outcome, or end of reaction

protons have a + charge 1 amu p^+ in nucleus
neutrons have no charge 1 amu n° in nucleus
electrons have a – charge 0 amu e^- in orbitals

Atomic mass = protons + neutrons = ___ amu

Mn 55 atomic mass
– 25 protons/electrons
30 = # of neutrons

Br 80 mass
–35 protons/electrons
45 = neutrons

Development of the Model of the Atom

Democritus— came up with the word and idea of atom, the smallest particle

Dalton – Billiard Balls—atoms hard spheres, only differences due to different masses of atoms by type of atom

Thompson – Plum Pudding Model (Cathode Ray Tube), discovered electrons, which he imagined were stuck in the positive charge of the atom "stuff"

Ernest J. Rutherford – Gold Foil Experiment– put electrons flying around a positively charged, dense nucleus

Niels Bohr – put the electrons into orbits, or energy levels; developed spectra

Modern Wave Mechanical Model—the electrons are understood to be in orbitals, or zones, not specific orbits

Counting atoms in a compound
Ethanol – C_2H_5OH = 9 atoms (2+5+1+1=9)
Magnesium hydroxide—$Mg(OH)_2$ = 5 atoms (1+ (2 X 2) =5)

Metalloid – transition substance between a metal and non-metal, included on the metalloid line except for (Al + Po)
Include : Boron, Silicon, Germanium, Arsenic, Antimony, Tellurium, Astatine

Isotope – chemically identical atoms with different numbers of neutrons, different masses

Zinc has 5 main isotopes
Zn 64 (48.89%) = 31.2896
Zn 66 (28.81%) = 19.0146
Zn 67 (4.11%) = 2.7537
Zn 68 (17.57%) = 11.9476
Zn 70 (0.62%) = .434 = 65.4395 = 65.4 amu

Electron configurations Ground state Neon – 2-8 Excited state Neon – 2-7-1

Noble Gases – have perfectly filled outer electron orbitals

Non-metals gain electrons to become – charged ions called Anions –1, –2, –3

Cation + Anion = neutral ionic compound

Group 1	Li	2-1	lose e^-	Li^{+1}
Group 2	Mg	2-8-2	lose 2 e^-	Mg^{+2}
Group 17	Cl	2-8-7	add e^-	Cl^{-1}
Group 16	S	2-8-6	add 2 e^-	S^{-2}
Group 15	N	2-5	add 3 e^-	N^{-3}

Cations form exactly at the same time as an anion

Cation	Anion	Neutral Compound	Name
Mg^{+2}	O^{-2}	MgO	Magnesium Oxide
Li^{+1}	SO_4^{-2}	Li_2SO_4	Lithium Sulfate
Ba^{+2}	$C_2H_3O_2^{-1}$	$Ba(C_2H_3O_2)_2$	Barium Acetate
K^{+1}	PO_4^{-3}	K_3PO_4	Potassium Phosphate

Molecular compounds share valence electrons

Prefixes are used to name molecular compounds

1 mono 2 di 3 tri 4 tetra 5 penta 6 hexa
7 hepta 8 octa 9 nona 10 deca

I_4O_9 = Tetra Iodine Nonoxide

Oxidation Numbers

Carbon	Phosphorous	compound formed from these oxidation numbers
-4	-3	C_3P_2 tricarbon diphosphide
+2	+3	C_3P_4 tricarbon tetraphosphide
+4	+5	C_5P_4 pentacarbon tetraphosphide

Cations	Anions	compound formed	Name
Ti^{+2} II	F^{-1}	TiF_2	Titanium (II) Fluoride
Ti^{+3} III	F^{-1}	TiF_3	Titanium (III) Fluoride
Ti^{+4} IV	F^{-1}	TiF_4	Titanium (IV) Fluoride

Mn_2O_7 = Manganese (VII) Oxide

Nickel (III) Dichromate = $Ni_2(CrO_7)_3$

Br_2S_6 = Dibromine hexasulfide

HONClBrIF Twins— elements that form stable only as pairs when pure

H	O	N	Cl	Br	I	F	
H_2	O_2	N_2	Cl_2	Br_2	I_2	F_2	= more stable as pairs

Allotropes – pure elements with different structures examples include oxygen + ozone ($O_2 + O_3$), or DIAMONDS and Pencil Graphite both pure carbon

Naming compounds
ionic – name the metal first, the non-metal second, change ending to –ide ammonium chloride is an ionic compound

Molecular — say the first name if a single atom, or else use a prefix. Leave end of name alone.
For the seconds name, use prefixes always, change to end with –ide. Ex CO, CO_2, H_2O

Table E ions keep their names

Transitional metal ionic compounds usually need Roman Numerals, if they make multiple cations

Ionic compounds	Molecular compounds
Metal + non-metal metals lose e^- to form cations + charge non-metals gain e^- to form anions – charge there is a transfer of electrons $\#e^-$ lost by cations must = $\#e^-$ gained by anions cation + anion charges sets the ratio of ions in compound all ionic compounds must be neutral	only form from non-metals no transfer of electrons no ions, no charges atoms share electrons in covalent bonds name – use prefix method set the ratios of atom:atom by the oxidation numbers

Group 1 = Alkali metals
Group 2 = Alkaline Earth Metals
Groups 3-12 = Transitional metals
Bottom 2 rows of table = Inner Transitional metals (all in group 3, periods 6 + 7)
Group 18 = Noble Gases
Group 17 = Halogens

1 mole = 6.02×10^{23} (Avogadro's Numbers)
½ mole = 3.01×10^{23}
2 moles = $12.04 \times 10^{23} = 1.204 \times 10^{24}$

$\frac{\text{0.5 moles Zn}}{1}$ X $\frac{6.02 \times 10^{23} \text{ atoms Zn}}{\text{1 mole}}$ = 3.01×10^{23} atoms of zinc

molar mass = mass of 1 mole of any substance

molar mass of magnesium oxide

<u>MgO</u>

Mg	1 X 24	
O	1 X 16	
	40 g/mole	

molar mass of Aluminum Permanganate

Al^{+3} MnO_4^{-1}

$Al(MnO_4)_3$

Al	1 X 27	27
Mn	3 X 55	165
O	12 X 16	192
		384 g/mole

How many moles are present here?

6.02×10^{23} molecules of Bromine

6.02×10^{23} molecules of Br_2 X $\frac{1 \text{ mole } Br_2}{6.02 \times 10^{23} \text{ molecules } Br_2}$ = 1 mole of bromine

4.81×10^{24} atoms Li X $\frac{\text{1 mole Li}}{6.02 \times 10^{23} \text{atoms Li}}$ = .799 = 7.99×10^{1} = 7.99×10^{1} moles

allotrope = example is "Bucky Balls" or C_{60}— different forms of purely one atom, also ozone and oxygen

spectra – line colors – energy given off when excited electrons in higher orbitals than normal return to their ground states, releasing, or emitting the energy they had previously gained.

<u>% composition by mass</u>

H_2O

H	1 X 2	$\frac{2g}{18g}$ X 100 = 11.1%	11.1 + 88.9 = 100%
O	1 X 16	$\frac{16g}{18g}$ X 100 = 88.9%	
	18 g/mole		

You have 3098 grams $HgCl_2$, name compound then determine how much of this mass is chlorine?

$HgCl_2$

Hg	1 X 201	201	$\frac{70}{201}$ X 100% = 34.8% chlorine in Mercury II Chloride
Cl	2 X 35	70	
		271g/mole	

Volume of a mole of GAS at STP is 22.4 liters   1 mole = 22.4 liters

Particles in 1 mole = 6.02×10^{23} Particles

Mass of one mole is = molar mass

If you have 181 L. of Argon, how many grams does it weigh? How many particles (atoms) of argon are present?

$181 \text{ L. Ar} \times \frac{1 \text{ mole}}{22.4 \text{ L.}} = 8.08 \text{ moles Argon}$

$\frac{8.08 \text{ moles Ar}}{1} \times \frac{40 \text{ g Ar}}{1 \text{ mole}} = 323 \text{ grams Ar}$

$8.08 \text{ moles Ar} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mole}} = 48.6 \times 10^{23}$ which equals 4.86×10^{24} atoms Ar

If you have 437 liters of Nitrogen at STP, what is its mass? How many particles (molecules) of N_2 are present?

Molar mass is 28g/mole

$437 \text{ L} \times \frac{1 \text{ mole}}{22.4 \text{ L}} = 19.5 \text{ moles}$

$19.5 \text{ moles } N_2 \times \frac{6.02 \times 10^{23}}{1 \text{ mole}} = 117.39 \times 10^{23} \rightarrow 117 \times 10^{23}$ adjusting the coefficient: 1.17×10^{25}

$19.5 \text{ moles } N_2 \times \frac{28\text{g}}{1 \text{ mole}} = 546\text{g } N_2$

Empirical Formulas: ratios of atoms in a compound, not the real formula, just the ratio in lowest or reduced form

$C_6H_6 = CH$
$C_5H_{10}O_5 = CH_2O$
$H_2O_2 = HO$
$H_2O = H_2O$
$C_{10}H_{22} = C_5H_{11}$

MOST EMPIRICAL FORMULAS ARE NOT REAL COMPOUNDS

Synthesis Reaction (AKA combination reaction): 2 or more smaller substances chemically combine into a bigger substance

Sodium and Chlorine form table salt
Skeleton – $Na + Cl_2 = NaCl$ Balanced – $2Na + Cl_2 = 2NaCl$

Iron + Oxygen form rust (Fe_2O_3)
Skeleton – $Fe_2 + O_3 = Fe_2O_3$ Balanced – $4Fe_2 + 3O_3 = 2Fe_2O_3$

Decomposition – larger molecules break down into smaller substances

Lead II oxide decomposes $2PbO = 2Pb + O_2$

Hydrogen monochloride decomposes $2HCl = H_2 + Cl_2$

Single Replacement – one atom will replace one ion in a solution

Always = Atom + cation/anion pair (which is an aqueous solution)

Zinc + Hydrochloric acid yields zinc chloride + hydrogen $Zn_{(S)} + 2HCl_{(AQ)} \rightarrow ZnCl_{2(AQ)} + H_{2(G)}$

Chlorine + Sodium bromide yields sodium chloride and bromine $Cl_{2(G)} + 2NaBr_{(AQ)} \rightarrow 2NaCl_{(AQ)} + Br_{2(G)}$

Double Replacement – start with 2 aqueous solutions should end up with one aqueous solution and one precipitate

Always = $cation/anion_{(AQ)}$ + $cation/anion_{(AQ)}$ = cation/ different $anion_{(AQ)}$ + cation/different anion $_{(S)}$

$3Na_2CO_{3(AQ)} + Mg_3(PO_4)_{2(AQ)} \rightarrow 2Na_3PO_{4(AQ)} + 3MgCO_{3(S)}$

$4KOH_{(AQ)} + Pb(NO_3)_{4(AQ)} \rightarrow Pb(OH)_{4(S)} + 4KNO_{3(AQ)}$

Combustion – always requires a hydrocarbon to combine rapidly to oxygen , forming $CO_2 + H_2O$ only.

Methane combusts: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Propane combusts: $C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2$

A hydrocarbon is a molecule that only contains carbon and hydrogen

Stoichiometry

You have 60.0 moles of oxygen, how many moles of water can form in this reaction?

$2H_{(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

Mole ratio $\frac{O}{H_2O}$ $\frac{1}{2}$ X $\frac{60}{x}$ x = 120 moles of water

Zinc reacts with hydrochloric acid. If 66 g zinc reacts completely, how many liters of gas form?

$Zn_{(AS)} + 2HCl \rightarrow ZnCl_{2(AQ)} + H_{2(G)}$

Mole ratio $\frac{Zn}{H_2}$ $\frac{1}{1}$ $\frac{1.02}{x}$ x = 1.02 moles $H_{2(g)}$

66g Zn = $\frac{1 \text{ mole}}{65 \text{ g}}$ = 1.02 moles

You have 1 lb. (454g) of Zinc that completely reacts in hydrochloric acid $HCl_{(AQ)}$ How many F.U.'s form.

$Zn_{(S)} + 2HCl_{(AQ)} = ZnCl_{2(AQ)} + H_{2(G)}$

$\frac{Zn}{ZnCl}$ $\frac{1}{1}$ $\frac{6.98}{X}$

454g Zn X $\frac{1 \text{ mole}}{65 \text{ g Zn}}$ = 6.98 moles Zn

6.98 moles X $\frac{6.02 \text{ X } 10^{23}}{1 \text{ mole}}$ = 42.0 X 10^{23} adjust the coefficient to 4.20 X 10^{24} F.U.'s $ZnCl_2$

Phase of matter	Solid	Liquid	Gas
Relative Kinetic Energy	Lowest	Middle	Highest

Unit	Standard Pressure
Atmospheres	1.0 atm
Millimeters of Mercury (Hg)	760 mm Hg
KiloPascals	101.3 kPa
Pounds per square inch (PSI)	14.7 psi

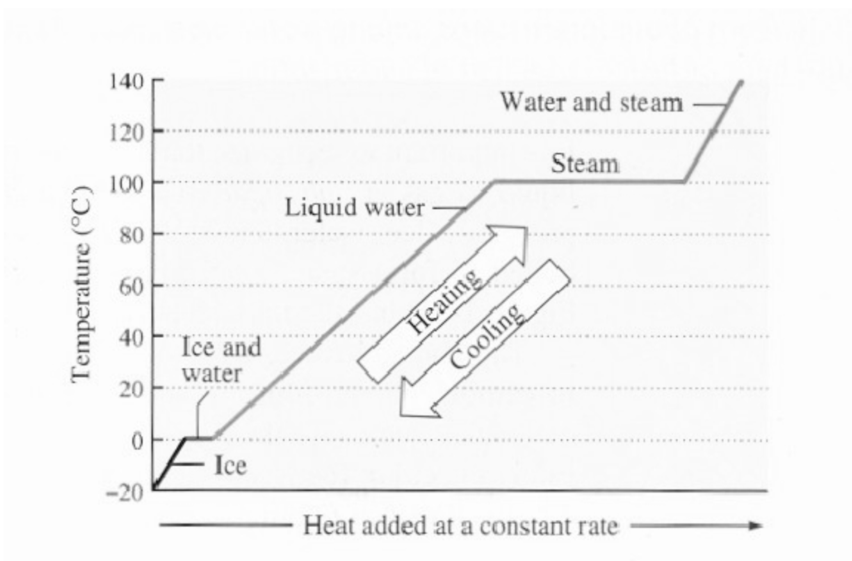
A gas is at the pressure of 1.25 atm. What is the pressure in kPa?

$$\frac{1.25 \text{ atm}}{1} \times \frac{101.3 \text{ kPa}}{1 \text{ atm}} = 126.625 = 127 \text{ kPa}$$

Your Location	Air Pressure	Boiling point of water
Below sea level (high pressure)	MORE than 101.3kPa	MORE than 100 degrees
At sea level (normal pressure)	101.3 kPa	100 degrees
Above sea level (lower pressure)	LESS than 101.3 kPa	LESS than 100 degrees

Intermolecular forces attract molecules of liquid together.
Evaporation = cooling process
Temperature is the measure of kinetic energy

Heating Curve and Cooling Curve



250. grams of ice at 273 K is warmed to water at 22°C. How much energy did that take?

warm up the ice to the melting point	melt the ice	add both together
$q = mH_F$ $q = 250$ grams (334 J/g) $q = 83{,}500$ J	$q = mC\Delta T$ $q =$ (250 grams)(4.18 J/g x °C)(22K) $q = 22{,}990$ J	106,490 Joules 106,000 J with 3 SF

Dynamic Equilibrium- always changing, always staying the same.

- Boiling is a function of kinetic energy and pressure
- Ionic solids have the highest melting points, molecular compounds have much lower melting points
- Solids have rigid, set, geometric pattern of particles
- When changing phases the kinetic energy stays steady, but there is a change in potential energy

Thermo-chemistry

1 cal. = 4.18 J   1 KJ = 1000 J   1000 cal. = 1 Calorie

$\Delta H-$ exothermic- gives off heat $\Delta H+$ endothermic – absorbs heat

Convert 286 cal. to Calories, kilojoules and joules

$$286 \text{ cal.} \times \frac{1 \text{ Calorie}}{1000 \text{ cal.}} = .286 \text{ Calories}$$

$$286 \text{ cal.} \times \frac{4.18 \text{ J}}{1 \text{ cal.}} = 1190 \text{ Joules}$$

$$1190 \text{ J} \times \frac{1 \text{ KJ}}{1000 \text{ J}} = 1.19 \text{ KJ}$$

You have 55 g H_2O at 15 ^{o}C . You heat it up to 17 ^{o}C. How much energy was absorbed?

$q = mC\Delta T$

$q = (55g)(4.18 \text{ J/g} \cdot \text{K})(2.0 \text{ K})$
$q = 460 \text{ J}$

When a cold pack is placed on an injured leg the heat moves from the leg to the cold pack

a bomb calorimeter measures the heat loss from food samples—used to measure calories in foods

GASES

Very small, round sphere particles; Particle volumes insignificant, don't attract or repel themselves;
Elastic collisions – no loss of energy, move in fast, straight lines; Can be greatly compressed.
Average K.E. is directly proportional to the temp. (Kelvin); Energy is transferred between colliding particles

$\frac{\text{Temperature}}{\text{Volume}}$ $\frac{\text{Kinetic Energy}}{\text{Temperature}}$ and $\frac{\text{Temperature}}{\text{Pressure}}$ Are Directly Proportional

directly proportional means, as one changes, so does the other in the same way.

$\frac{\text{Pressure}}{\text{volume}}$ is inversely proportional, which means as one goes up, the other goes down

As pressure decreases volume increases, the pressure exerted by gas is caused by the # of collisions

PV = constant so, $P_1V_1 = P_2V_2$

A sample of Neon gas at 125 kPa has a volume of 3.75 L. What is its gas constant

$$PV = C$$
$$(125kPa)(3.75L) = C$$
$$469 \text{ kPa X L} = C$$

A sample of Argon at STP and volume of 22.4 L is changed to 250 kPa. What is the new volume?

$$P_1V_1 = P_2V_2$$
$$(101.3 \text{ kPa})(22.4L) = (250 \text{ kPa})(V_2)$$
$$V_2 = 9.08 \text{ L}$$

A sample of gas of 406 mL at STP is changed to 50.5 kPa and expanded to 551 mL. What is new temperature of this gas?

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \frac{(101.3kPa)(406ml)}{273 \text{ K}} = \frac{(50.5kPa)(551ml)}{T_2}$$

$$\frac{7596361.5}{41127.8} = \frac{41127.8\ (T_2)}{41127.8} = 184.7 \text{ K} = 185 \text{ K} \quad (3 \text{ sf})$$

Avogadro's Hypothesis:
EQUAL VOLUMES OF DIFFERENT GASES AT THE SAME TEMPERATURE & PRESSURE HAVE THE SAME # OF PARTICLES, and the same nmber of MOLES of particles.

Examples…	Argon	CO_2	O_2
Volume	11.2 L	22.4 L	22.4 L
Temp	0 °C	273 K	273 K
Pressure	1 atm	101.3 kPa	0.5 atm
# particles	$3.01 \text{ X } 10^{23}$	$6.02 \text{ X } 10^{23}$	$3.01 \text{ X } 10^{23}$

When REAL gases act the most like IDEAL gases— At High Temp and Low Pressure

When comparing 2 or more gases: the gases with smaller particles are the ones that act more ideal

Helium is the most Ideal real gas

7 Trends of the Periodic Table

1. Atomic mass – going down one group increases the mass
2. Atomic size – going down one group increases p.m. (radius)
3. Ion size – cations are smaller than atoms anions are bigger than atoms
4. Metallic/non-metallic – most metallic (Fr) least metallic (He)
5. Net Nuclear Charge – all atoms are neutral but the nuclei is positive
6. Electronegativity – tendency to gain e^- in a covalent bond.
   The unit KJ/mol – group trend decreasing period trend increasing
7. Don't Memorize: look at 4 atoms in a period or group, and write out the trend values, see how they change

Periodic Table

Potassium (42) has 19 total protons with a mass of 42 amu. 42—19 = 23 neutrons

Chemical reactivity of group 17 elements decreases as you go on

Atomic number is equal to the number of p+ which also always equals the number of electrons

Elements in a periodic table are listed in order of atomic number, NOT ATOMIC MASS

$S^{-2} > S$ Anions are bigger than their atoms,

Na+1 < Na Cations are SMALLER due to less orbitals

1st ionization energy levels decrease when atomic radius increases (moving down a group)

BONDING

Out of the following elements (C, N, S, Si), silicon would make the most polar bond with oxygen
because the greatest difference in electronegativity values between oxygen and any of those 4 is with silicon

Cation + Anion = neutral ionic compound

Na atom 2-8-1 Na^{+1} cation 2-8
Cl atom 2-8-7 Cl^{-1} anion 2-8-8

Ions of Calcium and Sulfur combine this way: $[Ca]^{+2}$ $[\ S\ ]^{-2}$ CaS in a 1:1 ratio
A real Lewis Dot diagram would have 8 dots around the "S"

Mg^{+2} is isoelectric to Ne
P^{-3} is isoelectric to Ar

Water H
Ö: H

Carbon Dioxide O::C::O

Hydrogen monochloride H:Cl

Potassium Chloride $[K]^{+1}[\ Cl\ ]^{-1}$ the chlorine here would have 8 dots around it

Carbon disulfide S::C::S

2 metals don not bond with each other
Alloy – mixture of metals & non-metals—or by melting 2 or more metals together (brass is made by melting Zn + Cu)

Alloy can be metal and nonmetal too: Iron + Carbon = Steel

Covalent bonding – 2 or more non-metals that share electrons– no metals make covalent bonds

Octet rule – most atoms require 8 e^- to fill the orbital H, He, Li & Be this doesn't apply— they're too small

RELATIVE OXIDATION NUMBERS

CO_2

+4 + –2 + –2 = zero

CrO_4^{-2}

oxygen 4 x –2 = –8
Chromium must be +6
For the whole polyatomic anion to = –2

NaOH 1 + (– 2) + 1 = zero

Beryllium Flouride = BeF_2 $[Be]^{+2}$ $[F]^{-1}$ $[F]^{-1}$ both "F" would have 8 dots around them inside the brackets.

S doesn't have a noble gas e^- configuration S^{-2} does have a noble gas e^- configuration

A Fluoride ion is larger in radius than a Fluorine atom

Based upon intermolecular forces called hydrogen bonds, of these: (He O_2 SCl_2 NH_3) NH_3 has the highest BP He has only electron dispersion attractions (weakest). Oxygen does too. Sulfur dichloride has dipole attractions, the second weakest attraction. Ammonia has hydrogen bonding (the strongest intermolecular attraction.)

If Barium loses $2e^-$ it becomes a positive cation and the radius decreases (loss of e^- also means loss of whole orbital)

When e^- are transferred from one atom to another it forms an ionic bond

When bonds form energy is released

$CO_{(g)}$ contains 1 coordinate covalent bond (a weird "bond" to give octets all around.)
Oxygen "lends" pair of electrons to bond

A greater difference in electronegativity creates a stronger intermolecular force

PROPERTIES OF WATER*

1. Low vapor pressure
2. High Boiling Point
3. Surface Tension
4. Ice is less dense than liquid water
5. Water has a high specific heat capacity
6. Water forms solutions
7. Hydration of ionic crystals - things like copper (II) sulfate PENTAHYDRATE

*These are caused by **hydrogen bonding**

How many grams of NH_3 dissolves in 40 ml H_2O at 90°C (use table G)

$$\frac{NH_3}{H_2O}\ \frac{10g}{100ml} = \frac{Xg}{40ml} \qquad \frac{100X}{100} = \frac{400}{100} \qquad X = 4\text{ g } NH_3$$

Solubility – how much stuff dissolves into a solution
Solute – thing being dissolved
Solvent – what it is being dissolved in (often is water—when water is solvent it forms an AQUEOUS solution)
Saturated – maximum solute in solvent (solute is always in whole # grams from table G (no tenths)
Unsaturated – a solution that is not full yet– can hold more solute
Supersaturated – holds more solute than it should at given temperature
Solvation – process of dissolving into a solution

(if you saturate a solution at a high temp when it cools some of it falls out of solution, unless it supersaturates and holds more solute than it should. Bumping will cause precipitation of this excess solute)

adding ions to water at Standard pressure increases boiling point and decreases freezing point

water is polar (dihydrogen monoxide)
oxygen gets e^- most of the time in this molecule because oxygen has a much higher electronegativity value than H

When a solution contains ions it conducts electricity (it is an electrolyte) such as salt water. $NaCl_{(AQ)}$ conducts electricity and is an electrolyte. $NaCl_{(S)}$ does not conduct electricity, it has NO loose, mobile ions, but it is an electrolyte because when put into water, it would conduct electricity.

Solutions that are non-electrolytes have no ions, such as sugar water. No loose mobile ions = NON electrolyte
If two solutions can mix into each other they are Miscible
If they cannot dissolve into each other they are Immiscible
Dilute – small amount of solute (a weak solution)
Concentrated – a lot of solute
Solutions are homogeneous (same throughout)

Water of hydration – water bonded to a variety of ionic compounds
Ex. Magnesium Sulfate Heptahydrate $MgSO_4{\cdot}7H_2O$ or Copper (II) Sulfate Pentahydrate $CuSO_4{\cdot}5H_2O$
water is "hydrogen bonded" onto the ionic compound

(AQ) means that the solvent is H_2O

LIKE DISSOLVES LIKE, solutions that are polar can only dissolve polar or ionic compounds.
Non polar solvents can only dissolve non-polar compounds.

- Non-polar dissolves into Non-polar Polar dissolves into Polar (there are exceptions on Table F)

Colligative Properties – physical properties (boiling point, freezing point, vapor pressure)
These properties change if the solvent has compounds dissolved in them compared to the solvent alone
Salt = change in all 3 properties = slow evaporation, lower FP, increase BP, decrease VP
reason is that the water is attracted to itself, but now also to the ions, or polar compounds dissolved into water.

$NaCl_{(S)}$ goes into water → $Na^{+1}{}_{(AQ)} + Cl^{-1}{}_{(AQ)}$ = dissociation or ionization. Not a reaction, a phase change: S → AQ

How many grams KCl are in 1.75 M 2.00 L solution of KCl (aq)

$$\text{Molarity} = \frac{\text{\# of moles}}{\text{\# of liters}} \qquad 1.75\text{ M} = \frac{X \text{ moles}}{2.00\text{ L}} = 3.50 \text{ moles}$$

KCl
K 1 X 39 = 39
Cl 1 X 35 = 35
74g/mole

$$3.50 \text{ moles} \times \frac{74\text{g}}{1 \text{ mole}} = 259\text{g KCl}$$

How many grams $MgCl_2$ in a 3.50 M, 2.00 L solution of $MgCl_{2(aq)}$?

$$M = \frac{\text{\# of moles}}{\text{\# of liters}} \qquad 3.50 = \frac{X \text{ moles}}{2.00} = 7 \text{ moles}$$

$MgCl_2$
Mg 1 X 24 = 24
Cl 2 X 35 = 70
94g/mole

$$7 \text{ moles} \times \frac{94 \text{ g}}{1 \text{ mole}} = 658 \text{ grams } MgCl_2$$

You have 3.50 M $MgCl_{2(aq)}$. How do you make 1.75 M $MgCl_{2(aq)}$ of 78.0 ml?

$M_1V_1 = M_2V_2$
$(3.50M)(V_1) = (1.75M)(78.0ml)$
$V_1 = 39.0$ ml Stock

39.0 ml stock
+ sufficient water to get to 78.0 total mL
makes 78.0 mL 1.75 M $MgCl_{2(AQ)}$

You put 0.000050 g KCl into 2.000 L H_2O. What is the concentration in PPM?

$$PPM = \frac{\text{grams solute}}{\text{grams solution}} \times 1{,}000{,}000$$

$$2.00\ L \times \frac{1000\ ml}{1\ liter} = 2{,}000\ ml = 2{,}000.\ g\ water$$

$$\frac{0.000050\ g}{2{,}000.\ g} \times 1{,}000{,}000 = .025\ PPM$$

Colligative Property
Adding 1 MOLE of particles will change: Boiling point elevation = + 0.5 K
and the Freezing Point depression = –1.86 K

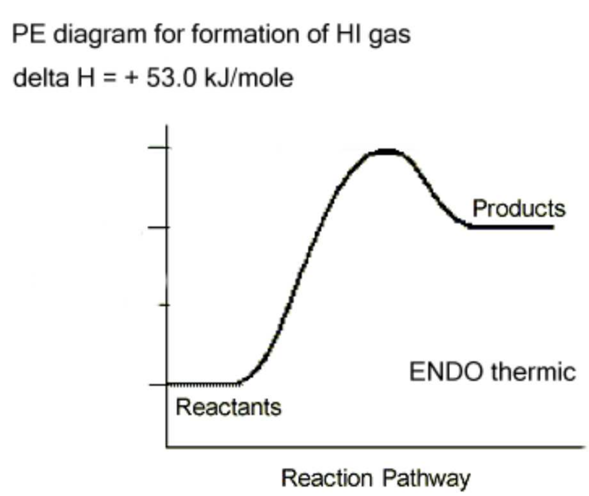
You put 94 gm $CaCl_2$ into 1.0 L H_2O, what is the new boiling point and freezing point
94 grams of $CaCl_2$ = 1 mole of $CaCl_2$, but in solution the compound breaks into 3 ions, so you get 3 moles of particles
.5 + .5 + .5 = 1.5 = 100 + 1.5 change in BP temp = 101.5°C for new BP

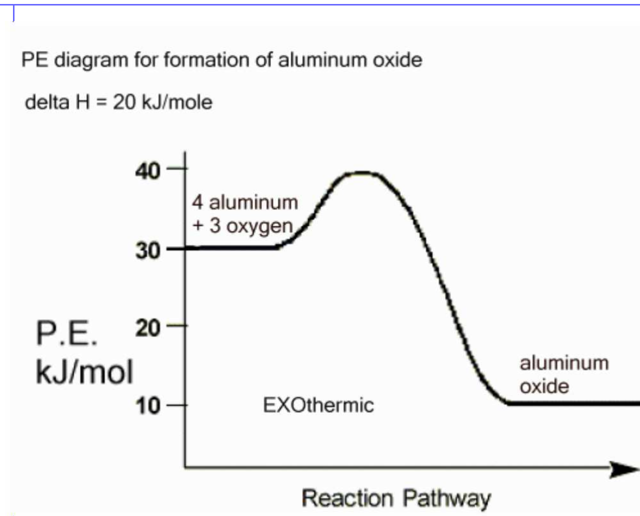
$CaCl_2$ = 3 moles of particles, so FP depression is (–1.86) + (–1.86) + (–1.86) = –5.58°C, so new FP = –5.58°C

You have a 100. ml solution of KI @ 293 K, what is the molarity?
145g X 1 mole = .873 moles

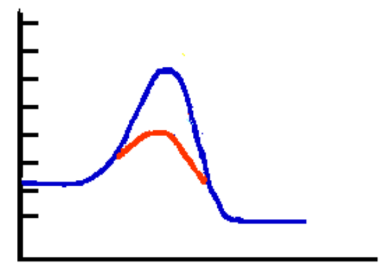
$$M = \frac{\#\ of\ moles}{\#\ of\ liters} = \frac{.873}{.100} = 8.73\ M$$

POTENTIAL ENERGY DIAGRAMS endothermic on the left, exothermic on the right
Measure of energy in reaction is in kilojoules/mole (kj/m)





Exothermic PE diagra



Exothermic potential energy diagram in blue, with Catalyst effect in red.

Catalysts lower the activation energy, providing an alternate pathway for the reaction to complete.

A catalyst does not change PE of reactants, PE of products or the ΔH

4 factors that affect the rate of chemical reactions

1. Temp — 1,2,3 increase the # of collisions
2. Surface area
3. Concentration
4. Catalyst — catalysts lowers the AE, by providing an alternate pathway forward

LeChatlier's Principle – when a chemical system is in dynamic equilibrium and a stress is added to this dynamic equilibrium the system will change to relieve the stress, and create a new dynamic equilibrium

$$N_2 + 3H_2 \rightleftharpoons 2NH_3 + \text{energy}$$

Forward reaction →

← reverse reaction

Forward reaction is synthesis (exothermic) Reverse reaction is decomposition (endothermic)

$$2SO_{2(G)} + O_{2(G)} \rightleftharpoons 2SO_{3(G)} + \text{energy (37.2 kj)}$$

these stresses added

will result in equilibrium shifts of this direction until a new equilibrium forms

Add SO_2 →

← Add SO_3

← Remove O_2

Remove SO_2 →

Add O_2 →

← Increase Temp.

Increase Pressure →

Decrease Temp. →

Add Catalyst NEITHER, just get to equilibrium faster

Pressure only affects the gases, just count the moles on each side of the arrows in the equilibrium

$$2C_{(S)} + 3H_{2(G)} \rightleftharpoons C_2H_{6(G)} + 84.0 \text{ KJ}$$

← Increase Temp

Increase Pressure → 3 moles: 1 mole

Remove Ethane →

← Remove Carbon

Decrease Temp. →

← Decrease Pressure 3 moles: 1 mole

Entropy – measure of disorder in a system
lower temp. Means less entropy gases have the highest entropy of all three phases
high temp. means more entropy solids have the lowest entropy of all three phases
big molecules have less entropy than small molecules

ACIDS/BASES

Arrhenius Theory

Acid – aqueous solution that contains excess $H^{+1}_{(AQ)}$ hydrogen cations

Base – aqueous solution that contains excess $OH^{-1}_{(AQ)}$ hydroxide ions

Acid + Base = Salt + Water (a salt in any ionic compound, metal cation + nonmetal anion)

(examples)…
$H_3PO_4 + 3KOH \rightarrow K_3PO_4 + 3HOH$
$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2HOH$
$HCl + KOH \rightarrow KCl + HOH$
$HC_2H_3O_2 + 2NaOH \rightarrow Na_2C_2H_3O_2 + HOH$

Alternate, or Bronsted Lowry Theory of Acids and Bases (we use for explaining ammonia as a base)

Base – accepts H^+
Acid – donates H^+
$NH_{3(G)} + HOH_{(L)} = NH_4^{+1}{}_{(AQ)} + OH^-{}_{(AQ)}$

NH_3 becomes NH_4^{+1} by accepting the H^+

HOH donates the H^+ and becomes the OH^-

A neutral pH is when the pH = 7 which happens when the $H^+ = OH^-$
or there are no $H^+_{(AQ)}$ or $OH^-_{(AQ)}$ is said to be neutral

pH is the measure of H^{+1} in solution with a negative log scale, which runs from 0 to 14,
lower than 7 = acid; greater than 7 = base
Each whole number change in pH is equal to a 10X change in hydrogen ion concentration or strength

a solution of pH 5.0 is 10X more acidic (has 10X more H^+ ions) than a solution of pH 6.0

A solution of pH 4.0 is 100X more acidic (100X more H^+ ions) than a solution of pH 6.0

A solution of pH 13.0 is 10,000X more basic (10,000 X more OH^- ions) than a solution of pH 9.0

When 50. milliliters of HNO_3 solution is neutralized by 150 milliliters of a 0.50 M solution of KOH, what is the concentration of HNO_3?

$$(\#H^{+1})M_AV_A = M_BV_B(\#OH^{-1})$$
$$(1)(M_A)(50.ml) = (.50M)(150ml)(1)$$
$$M_A = 1.5\ M$$

If you need 24.2 ml of 1.00M HCl to exactly neutralize your 37.0 ml of NaOH, what is the concentration of the base?

$$(\#H^{+1})M_AV_A = M_BV_B(\#OH^{-1})$$
$$(1)(1.00)(24.2) = (M_B)(37.0)(1)$$
$$M_B = .654M$$

If you need 45.7 mL of 2.36M $H_2SO_{4(AQ)}$ to neutralize 344 mL of NaOH, what is base molarity?

$$(\#H^{+1})M_AV_A = M_BV_B(\#OH^{-1})$$
$$(2)(2.36)(45.7) = (M_A)(344)(1)$$
$$215.704/344 = M_A$$
$$0.627\ M = M_A$$

An Acid can be described these ways: $H_3O^{+1}{}_{(aq)}$ = hydronium ion $H^{+1} + H_2O_{(L)} = H_3O^{+1}{}_{(aq)}$

or as just an AQ H^{+1} or as a proton in water, or as any substance that can donate a H^{+1}

Titration = mixing acid + base to neutralize it (using burets and indicator solutions)

A chemical formula for the negative ion in aqueous nitric acid solution is NO_3^{-1}

$CaCO_3 + 2HCl = \underline{CaCl_2} + H_2O + CO_2$ $CaCl_2$ is aqueous

Organic Chemistry

Carbon – central atom of O.C.

Meth = 1 carbon
ane = all single bonds between C and H in molecules example: CH_4

Homologous - same body or root
Series - set or group of similar stuff
Hydrocarbons - organic molecules made of carbon + hydrogen

Alkanes – chains of carbon atoms only single bonds C-C Generic formula – $C_nH_{2n+2} = CH_4$

```
                H H
2 carbon alkane H-C-C-H    = ethane
                H  H
```

C_2H_6

7 carbon alkane -C-C-C-C-C-C-C- = heptane

Nonane $CH_3(CH_2)_7CH_3$ Decane $CH_3(CH_2)_8CH_3$

Alkenes – chained hydrocarbons with at least one C=C double bond

Propene –C=C-C– 2-Butene –C-C=C-C–

Alkynes – hydrocarbon chain with at least 1 C≡C triple bond

```
       Cl Br
Cl–C≡C–C–C–C–   5 bromo; 1,4,4 trichloro; 3 flouro 1· pentyne
       F  Cl
```

Substitution reactions with Halogens substituting one atom of a halogen to an alkane. Can't add both halogens, ther e is, no room in the molecule, so only one atom of diatomic HONClBrIF twin subs in. .

Methane $CH_4 + Cl_2 = CH_3Cl + HCl$

Chloromethane $CH_3Cl + Cl_2 = CH_2Cl_2 + HCl$

Alcohols R-OH (molecular and dissolves into water)

Ethanol -C-C-OH 1·Hexanol -C-C-C-C-C-C-OH

Organic Acids

R-C–OH (with =O on the C)

Butanoic Acid -C-C-C-C-OH (with =O on the last C)

Esters R-C-O-R' (with =O on the C) methyl propanoate -C-C-C-O-C (with =O on the third C) name the "oxygen tail" first

Amines R—N-H (with H on the N) in our class the both "R" groups attached to the nitrogen will be H, or else we can't name them with our knowledge.

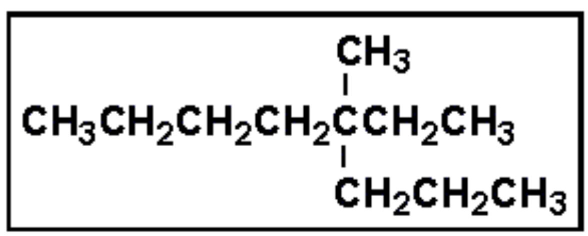
1·Propanamine -C-C-C-N-H_2 = $CH_3CH_2CH_2NH_2$
this is "number 1" because the amine group is attached to carbon #1

Amides R—C—N–H (with =O on the C and H on the N)

Propanamide -C-C-C—N–H_2 (with =O on the third C) this needs no number, it always has to be at the end of chain

Branched Chains

At right is an 8 carbon chain, so it's LAST name is octane.
This is 4ethyl, 4 methyl octane



Substitution – 1:1 exchange hydrogen for halogen Ex.(ethane + chlorine) -C-C- + Cl_2 = -C-C-Cl + HCl

Addition – unsaturated hydrocarbon and halogen Ex. (ethene + Flourine) -C=C- + F_2 = F-C-C-F

Esterfication – organic acid + alcohol = ester + HOH Ex. (propanoic acid + ethanol = ethyl propanoate + HOH

-C-C-C-OH + HO-C-C- = -C-C-C-O-C-C- + HOH (with =O on the third C of the acid and of the ester)

Polymerization – plastics Poly = many Mer = unit/body

Chloroethene CH_3CH_2Cl X an unknown number N of these molecules can be make to break open the double bond to allow for these
chloroethenes to link together forming long chains of polychloroethenes

Fermentation – anerobic respiration by yeast
Sugar along with yeast/enzymes = carbon dioxide and ethanol
$C_6H_{12}O_6$ (with yeast in water) = $2CO_{2(g)} + 2C_2H_5OH$

Saponification – making of SOAP

Isomers – chemically identical with different structures

Ethyl methyl ether and Propanol
-C—O—C-C- = -C-C-C—OH

C_3H_8O is the same chemical (or molecular) formula, but very different structures, different functional groups

Cyclo-hydrocarbons = ringed carbon chains

Cyclopropane ≠ propane $C_3H_6 \neq C_3H_8$

Fractional distillation is to purify a little at a time. Crude oil has mixture of many petroleum oils. Heat low temp and ONE might boil away. Distill it into one container. Heat up a bit more, a second oil reaches it's boiling point, collect that one. Keep going until you have boiled them all apart.

C_2H_4, C_3H_6, C_4H_8 all belong to the same homologous series

C=C = 2 pairs of shared e^- C≡C- = six electrons being shared in 3 covalent bonds

$CH_3CH_2CH_2CHO$ = Butanal

Ethers R-O-R (oxygen bridge molecule)	Aldehydes	Ketones
Butyl propyl ether -C-C-C-C—O—C-C-C-	O ‖ R-C-H at the end of a chain only propanal CH_3CH_2CHO O ‖ -C-C-C-H	O ‖ R-C-R are in the middle of a chain only 3·heptanone O ‖ -C-C-C-C-C-C-C-

REDOX

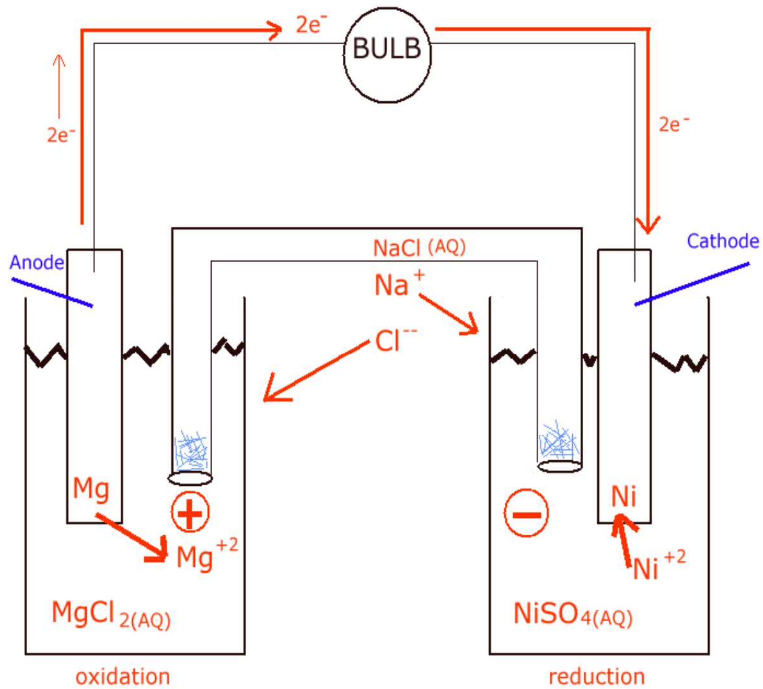
Oxidation – loss of electrons (LEO)

Reduction – happens at the Cathode (Red-Cat)

Reduction – gain of electrons (GER)

Oxidation – happens at the Anode

Voltaic Cell



½ OX $Mg^{o} = Mg^{+2} + 2e^{-}$

½ Red $Ni^{+2} + 2e^{-} = Ni^{o}$

3 reasons why batteries die

1. Run out of salt ions
2. Run out of anode metal
3. Run out of cathode cations

Example half reactions, then by combining them by canceling out the electrons...

½ OX $Ni^{o} = Ni^{+2} + 2e^{-}$

½ Red $2Ag^{+1} + 2e^{-} = 2Ag^{o}$

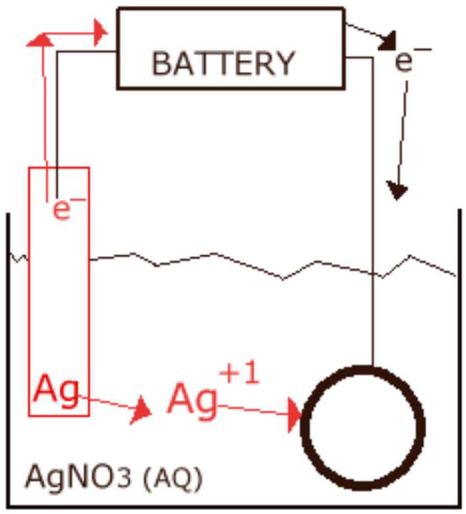
Net Ionic Equation $= Ni^{o} + 2Ag^{+1} = 2Ag^{o} + Ni^{+2}$

Electrolytic Cells

Putting a copper ring into silver nitrate solution… The spontaneous reaction that SHOULD occur here is
$Cu_{(s)} + AgNO_{3(aq)} = CuNO_{3(aq)} + Ag_{(s)}$

But the Electrolytic cell uses electricity to force redox… So this happens instead:

½ Ox $Ag^{o} = Ag^{+1} + 1e^{-}$ (silver bar into solution)
½ Red $Ag^{+1} + 1e^{-} = Ag^{o}$ (silver ions from solution onto ring)



In Voltaic cells: chemical energy spontaneously creates electricity.
In Electrolytic cells: outside energy is required (electricity usually) to force a redox reaction to occur.