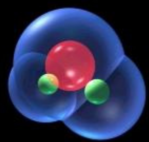
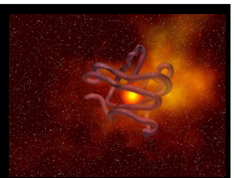
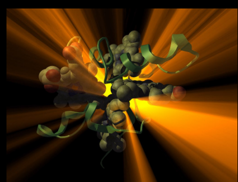
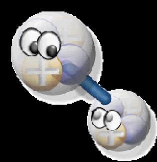




Bonding





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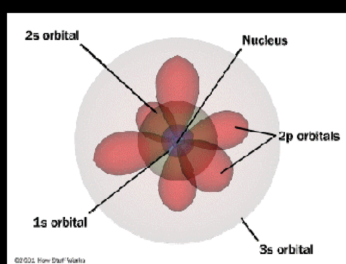
Bonding Involves Outer Orbitals



A matter of geometry – atoms only contact outer regions

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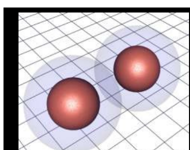

Chemical Bonds

Attractive force that holds 2 or more atoms together in a unit
Energy of bonded pair less than energy of separated atoms

Basic Types

Ionic:
Transfer of electrons from one atom ➔ ions
+/- ions attracted to one another
Strong electrostatic forces hold ions within crystal matrix

Covalent:
Sharing a pair of electrons between two nuclei

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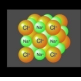
General Bond Properties

	Ionic	Covalent
Basic Component	Ions (Charged Matrix)	Atoms/Molecules
Constituents	Metal + Non-Metal	2 Non-Metals
State (RT)	Solid	Solid, Liquid, Gas
Melting Point	Very High (> 200 °C)	Lower (< 200 °C)
Odor	None	May Be Present
Flammability	No	May Be
Conductivity	Solids: Poor Melted: Good Aqueous: Good	Solids: Poor Melted: Poor Aqueous: Poor

In general:
Molecular compounds
Are Not Electrolytic, but some
Exceptions (acids) are known



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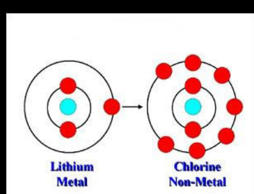
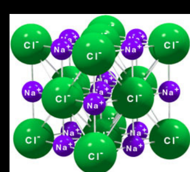
Ionic Interactions

Not a single entity between individual atoms …
Strong electrostatic forces hold ions within crystal matrix





Commonly, Metal Cation & Non-Metal Anion

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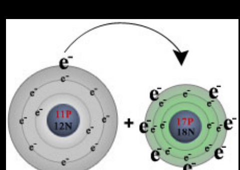
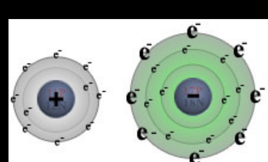
Ionic Interactions

Transfer of electrons from one atom to another to form ions





Both atoms have inert (filled outer shells) configuration
Cation smaller than neutral atom
Anion larger than neutral atom

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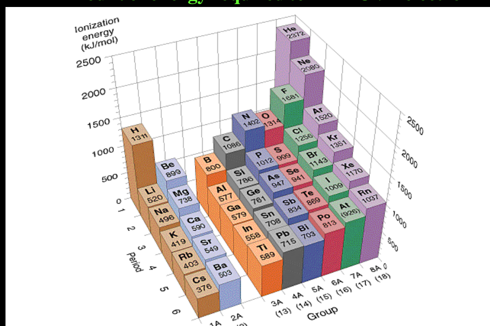
Ionization Energy

Amount of energy required to REMOVE electron



Most likely to form cation

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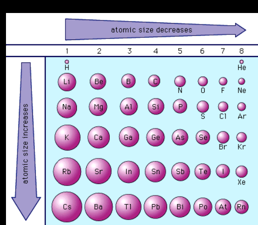
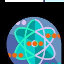
Atom Size

Lowest ionization energy→ e^- largest distance from nucleus



Same Group: (Moving Down)
- Each row has more particles
- Occupies more space

Same Period (Moving Across)
- Each column has more protons
- More positive charges
- More attraction to electrons
- Decrease in size



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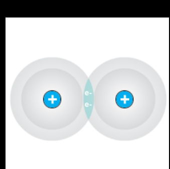
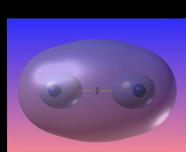
Covalent Bonding

Covalent = Sharing
Electron density shared between atoms





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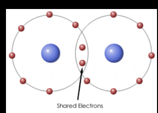
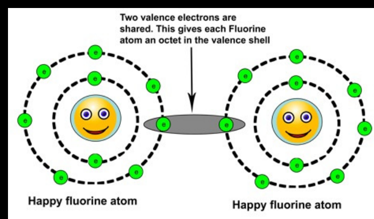
Covalent Bonding

Atoms share electrons to complete "Octet"





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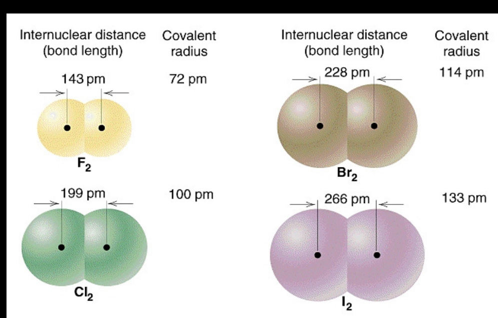
Nonpolar Covalent: equal sharing of e⁻



Atomic Radius Influences Bond Length (Strength)

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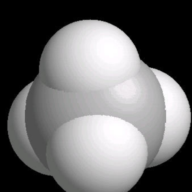
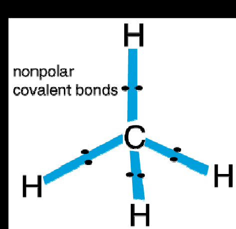
Non-Polar Covalent Bonding: Methane (CH_4)





Carbon & Hydrogen ~ Same Electronegativity

Equal Electron Sharing ➔ Non-Polar Covalent Bond

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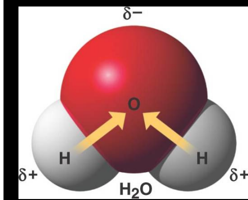
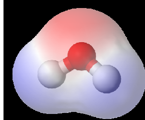
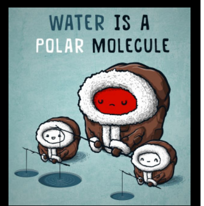
Polar Covalent Bond: Water (H_2O)







Oxygen More Electronegative than Hydrogen

Unequal Electron Sharing ➔ Polar Covalent Bond

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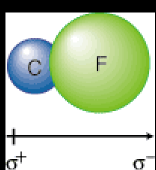
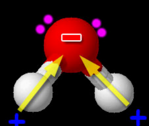
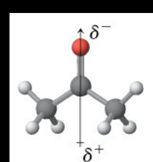
Dipole







Result of non-uniform distribution of electrons (charges)

Arrow ➔ drawn with arrow head at most negative

Direction reflects relative direction of charge separation

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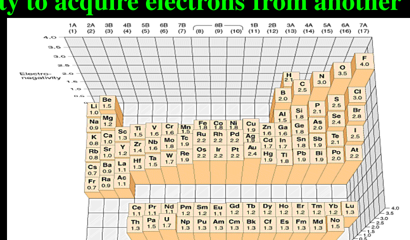
Electronegativity

Qualitative Measure of ability to acquire electrons from another atom



Applies only to single, isolated atom

Measured in Pauling Units

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Electronegativity Values

Electronegativity

1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A
							H 2.1									
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2

<1.0 1.5–1.9 2.5–2.9

1.0–1.4 2.0–2.4 3.0–4.0

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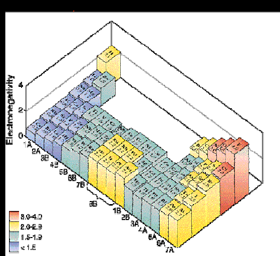
Electronegativity Differences

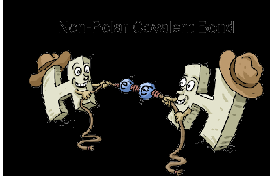
Δ ≤ 0.4 ➔ non-polar covalent

Δ > 0.4 - 1.9 ➔ polar covalent

Δ > 1.9 ➔ ionic

Δ = difference in electronegativity of the bonded





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Use Table of Electronegativities to determine bonds type

H–F
F = 4.0
H = 2.1
1.9 Polar-Covalent

Cl–F
F = 4.0
Cl = 3.0
1.0 Polar-Covalent

Na–F
F = 4.0
Na = 0.9
3.1 Ionic

Ca–F
F = 4.0
Ca = 0.7
3.3 Ionic

Δ ≤ 0.4 ➔ non-polar covalent

Δ 0.4 - 1.9 ➔ polar covalent

Δ > 1.9 ➔ ionic

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Indicate which is the more polar bond
Indicate the polarity of the dipole

C–O or Si–O H–O or H–S H–S or H–I H–P or H–S

> Electronegativity difference, > polarity
Most electronegative atom ➔ negative end of dipole

C–O or Si–O	H–O or H–S	H–S or H–I	H–P or H–S
C = 2.5	H = 2.1	H = 2.1	H = 2.1
O = 3.5	O = 3.5	S = 2.5	P = 2.1
Si = 1.8	S = 2.5	I = 2.5	S = 2.5
Δ CO = 1.0	Δ HO = 1.4	Δ HS = 0.4	Δ HP = 0.0
Δ SiO = 1.7	Δ HS = 0.4	Δ HI = 0.4	Δ HS = 0.4
SiO more polar	OH more polar	Same polarity	HS more polar
O is negative	O & S negative	S & I negative	S is negative

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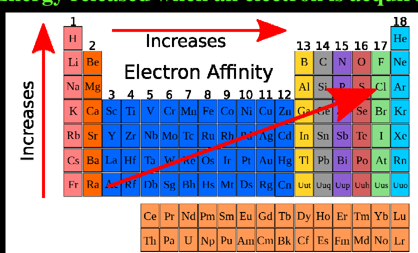
Electron Affinity

Quantitative Measure of
Energy released when an electron is acquired



Applies to either single atom or molecule
Measured in Kj/mol or eV

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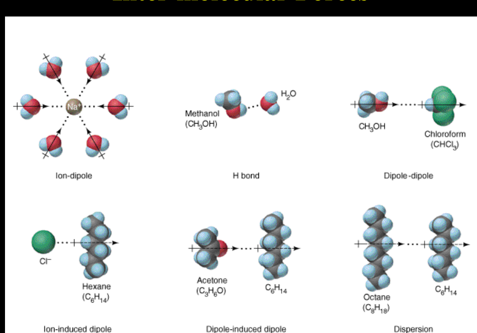
Inter-molecular Forces



Interactions Between Molecules

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Strength of Inter-molecular Forces

Individually weak
Abundant presence; changes physical properties
Defines shapes / interactions of molecules

Type of Interaction	Energy Range (kJ/mol)
Intermolecular	
Van der Waals	0.01 - 10
Hydrogen bond	10 - 40
Chemical Bond	
Ionic	100 - 1000
Covalent	100 - 1000

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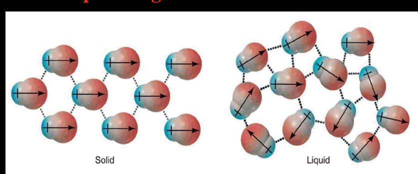
Dipole-Dipole Interactions

Molecules with permanent Dipole
Dipoles align … cohesive attraction



Alters physical properties
Typically increases melting/boiling point
energy needed to overcome multiple interactions
Water NH_3

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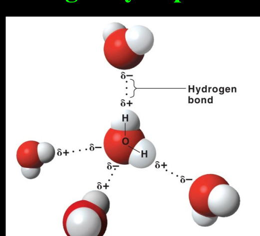
Hydrogen Bonds
A Biologically important dipole-dipole interaction



Low Energy (weak)
Individually weak,
But, significant in quantity

Pairs
H & Electronegative Atom
(especially N & O; F)

A strong dipole-dipole interaction

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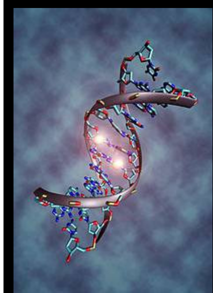
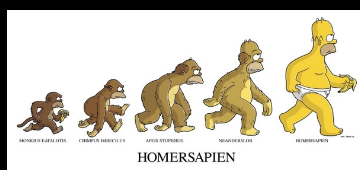
Hydrogen Bonds



Hydrogen Bonds define molecular shapes

Low energy of each bond allows DNA replication
Also makes the molecule susceptible to change

The fragility of DNA allows for:
mutation / evolution / cancer



DNA Double Helix

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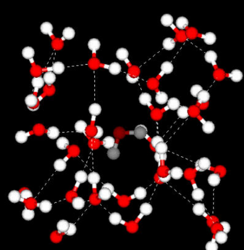
Hydrogen Bonds - Water



Boiling Point (°C)

H_2O = 100
H_2S = - 60.7
H_2Se = - 42
H_2Te = - 2
H_2Po = 37

If Mass Controlled,
H_2O should boil
At ~ -100 °C

Energy needed to overcome H-bonded network is considerable

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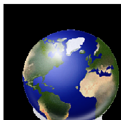
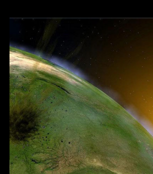


Water (H_2O)

If Water Boiled at - 100 °C,
Earth would have no liquid water
And
Life as we know it, would not exist



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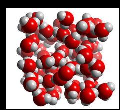
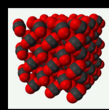
Water & Dry Ice

Water: Ice	"Dry Ice" (CO_2)
H-Bonded Network	No Bonding Network
Melting Point 0 °C	Sublimes (-78.5 °C)

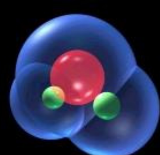
H-bonded network requires energy to disrupt

London Dispersion Forces (Van der Waals's)

Weakest interaction
(inversely proportional to r^6 between atoms)
Temporary; when adjacent atom electrons create dipole
All atoms; more prevalent in heavier/larger
Stronger when atoms easily polarized
At 3 Angstrom, ~ 1 kcal/mole

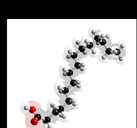
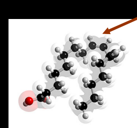
Van der Waal Radii
Volume of space where significant

Biologically (especially in lipids) significant



London Dispersion Forces in Fats

"Saturated" Fats are mostly linear molecules

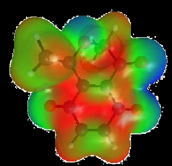
Site of Unsaturation
(a double bond)
Puts a "kink" in the otherwise, linear chain

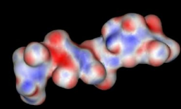
"Unsaturated" Fats are bent molecules

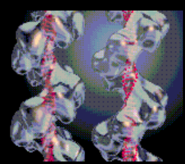
Saturated fats - linear molecules bundle together
this takes a lot of energy to undo (melt) ➔ solids
Unsaturated fats – "kinks" prevent bundling ➔ liquids

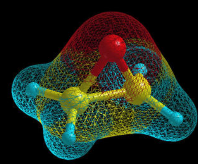
Van der Waal Radii Approximates Molecular Influence

Used to understand
molecular architecture
& how molecules interact









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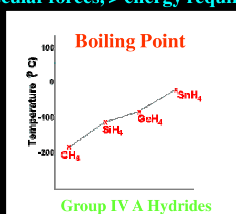
Physical Properties & Intermolecular Forces

Control Physical properties (State of Matter)
Melting & Boiling points
Result of progressive elimination of intermolecular forces
> intermolecular forces, > energy required to melt/boil



If only dispersion forces present (no H-bonding),
the more mass present (higher Z), > boiling point

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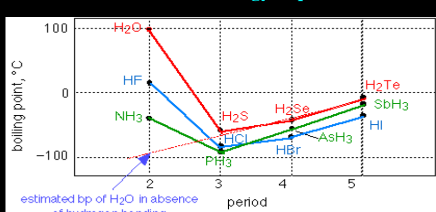
Physical Properties & Intermolecular Forces

Control Physical properties (State of Matter)
Melting & Boiling points
Result of progressive elimination of intermolecular forces
> intermolecular forces > energy required to melt/boil



If H-bonding present, H bonded b.p. higher & off-line

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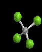
Solubility and Intermolecular Forces

Polar solutes dissolve in water (polar solvents)
Non-polar solutes dissolve in non-polar solvents

"Like Dissolves Like"



Solvent	Chemical Formula	Boiling point	Dielectric constant	Density
Non-Polar Solvents				
Hexane	$CH_3\cdot CH_2\cdot CH_2\cdot CH_2\cdot CH_2\cdot CH_3$	69 °C	2.0	0.655 g/ml
Benzene	C_6H_6	80 °C	2.3	0.879 g/ml
Toluene	$C_6H_5\cdot CH_3$	111 °C	2.4	0.867 g/ml
Diethyl ether	$CH_3CH_2\cdot O\cdot CH_2\cdot CH_3$	35 °C	4.3	0.713 g/ml
Chloroform	$CHCl_3$	61 °C	4.8	**1.498 g/ml**
Ethyl acetate	$CH_3\cdot C(=O)\cdot O\cdot CH_2\cdot CH_3$	77 °C	6.0	0.894 g/ml
Polar Aprotic Solvents				
1,4-Dioxane	$\underline{/\cdot CH_2\cdot CH_2\cdot O\cdot CH_2\cdot CH_2\cdot O\cdot\backslash}$	101 °C	2.3	**1.033 g/ml**
Tetrahydrofuran (THF)	$\underline{/\cdot CH_2\cdot CH_2\cdot O\cdot CH_2\cdot CH_2\cdot\backslash}$	66 °C	7.5	0.886 g/ml
Dichloromethane (DCM)	CH_2Cl_2	40 °C	9.1	**1.326 g/ml**
Acetone	$CH_3\cdot C(=O)\cdot CH_3$	56 °C	21	0.786 g/ml
Acetonitrile (MeCN)	$CH_3\cdot C\equiv N$	82 °C	37	0.786 g/ml
Dimethylformamide (DMF)	$H\cdot C(=O)N(CH_3)_2$	153 °C	38	0.944 g/ml
Dimethyl sulfoxide (DMSO)	$CH_3\cdot S(=O)\cdot CH_3$	189 °C	47	**1.092 g/ml**
Polar Protic Solvents				
Acetic acid	$CH_3\cdot C(=O)OH$	118 °C	6.2	**1.049 g/ml**
n-Butanol	$CH_3\cdot CH_2\cdot CH_2\cdot CH_2\cdot OH$	118 °C	18	0.810 g/ml
Isopropanol (IPA)	$CH_3\cdot CH(\cdot OH)\cdot CH_3$	82 °C	18	0.785 g/ml
n-Propanol	$CH_3\cdot CH_2\cdot CH_2\cdot OH$	97 °C	20	0.803 g/ml
Ethanol	$CH_3\cdot CH_2\cdot OH$	79 °C	24	0.789 g/ml
Methanol	$CH_3\cdot OH$	65 °C	33	0.791 g/ml
Formic acid	$H\cdot C(=O)OH$	100 °C	58	**1.21 g/ml**
Water	$H\cdot O\cdot H$	100 °C	80	1.000 g/ml

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Hydrophobic Materials "Repel" Water

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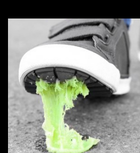
Chewing Gum Similar to Rubber Soles & Asphalt



Stuck Gum Removed with Oily Material

Oil Breaks Adhesion

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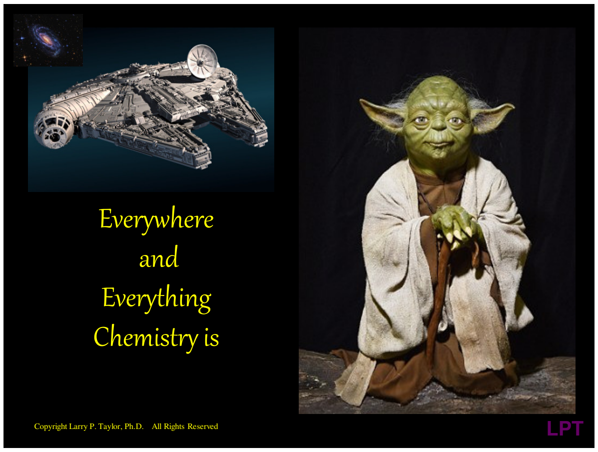
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