Block 3
F'16

Chemistry Unit 4

Primary reference: *CHEMISTRY*, Addison-Wesley

Topic	Essential Knowledge	Study Support
Scientific Investigation **1.4** **SOL 1a,b,f**	Use chemicals and equipment safely. **Accuracy** is how close a measurement is to the true value. An accurate measurement has very little error. Percent Error = 100 x | accepted value-exper. value | /accepted value	
Atomic Structure and Periodic Relationships **2.3** **SOL 2d, 2g, 2i**	**Niels Bohr** proposed the planetary model of the atom with electrons located in distinct energy levels (orbits) around the nucleus. **Louis de Broglie** proposed that all particles have wavelengths. (including electrons). **Max Planck** proved that a photon's wavelength is proportional to its energy. **Schrodinger** calculated the theoretical shapes of electron orbitals (s,p,d,f). **Heisenberg** developed the uncertainty principle concerning an electron's location and velocity. Electrons are added one at a time to the lowest energy levels first (**Aufbau Principle**). An orbital holds a maximum of two electrons (**Pauli Exclusion Principle**). Electrons occupy equal-energy orbitals so that a maximum number of unpaired electrons results (**Hund's Rule**). Energy levels are designated 1–7. Orbitals are designated s, p, d, and f according to their shapes(sphere, dumbbell, 4-leaf clover.) The s, p, d, f orbitals relate to regions of the Periodic Table. **Valence electrons** occupy the highest principle energy level of an atom. All the elements in a group have the same number of valence electrons. An element's electron configuration determines the number of valence electrons. Example: Bromine's valence electron configuration is $4s^24p^5$ with 7 valence electrons. The outermost electrons in an atom are called **valence electrons.** The period (row) number on the periodic table corresponds to the outermost energy level occupied by the valence electrons in an element. Elements in the same group (column) on the periodic table have the same number of valence electrons **Lewis dot diagrams** show the valence electrons of an atom. The electrons (dots) are arranged around the element's symbol. **Metallic bonds** consist of the attraction of free-floating valence electrons for the positively charged metal ions.	**Ch 13: Read pp 361-364** on development of modern quantum mechanical model. **Ch 13: Read pp 364-366** on orbital shapes. **Ch 13: Read pp 367-369** on electron configurations and Aufbua principle. **Ch 13: Read pp 372-381** on the relationship between atomic orbitals and atomic emission spectra. **Ch 15**: Read pp 413-424 on ionic compounds. Read pp 427 on metallic bonds.
Nomenclature, Formulas, and Reactions **3.3** **SOL 3a, 3d, 3e**	Bonds form between atoms to achieve stability. Ionic compounds are formed by the attraction between positive and negative ions. Ions are formed by **electron transfer** from a metal to a non-metal. After electron transfer, both ions meet the **octet rule**. The octet rule is the tendency of an atom to take on the configuration of a noble gas.	**Ch 6**: Review pp 146-156 for naming and writing formulas for ionic compounds.
Molar Relationships **4.4** **SOL 4d**	Dissolving is a physical change that involves heat. When an Ionic compound dissolves in water it breaks into the ions that make it up. This process is called **dissociation** and can be expressed by an equation. Example: $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ Ionic compounds that dissociate completely in water are strong electrolytes.	**Ch 17:** read pp 483-485 Due 11/1
Phases of Matter and Kinetic Molecular Theory **5.3** **SOL 5d,5e,5f**	**Specific Heat Capacity**(C) is a physical property of a substance. $\boxed{Q = mC\Delta T}$ is use to calculate heat, mass, specific heat or temperature change respectively. Specific heat can be used to identify a substance. Atoms and molecules are in constant motion. Forces of attraction between molecules determine the physical changes of state. The intermolecular forces must be overcome in order for a substance to melt or boil. Phase changes that require heat (like melting or boiling) are **endothermic.** ΔH is positive for an **endothermic** change. This means heat goes in. **Molar heats of fusion and vaporization** can be used to calculate energy changes. Phase changes that give off heat (like freezing and condensing) are **exothermic.** ΔH is negative. This means heat is released. **Molar heats of solidification and condensation** can be used to calculate energy changes. Heating and cooling curves, known as **temperature line graphs,** show the energy changes that occur as a substance goes from a solid to a gas as temperature is changed.	**Ch 11:** read pp 293-302 on heat capacity and specific heat capacity. ✓ Read pp 307-311 on molar heats of phase changes. Review the temperature line graph in figure 11.15 on page 310.

Objectives for Unit 4

Chemistry, Addison-Wesley, 2002

Topic Outline

I) Thermochemistry Part 1(Chapter 11)
  A) Types of Energy
  B) Exothermic and Endothermic processes
  C) Heat Capacity and Specific Heat (p 295-299)
    1) Calculations using specific heat capacities (p.299: 1-3, 8-10)
    2) Calorimetry (p. 300-306)
  D) Changes of State and Heat Changes (p. 307-313)
    1) Phase Changes and Interpreting Heating Curves
    2) Molar Heats of Fusion and Solidification (p.309: 20, 21)
    3) Molar Heats of Vaporization and Condensation (.311: 22, 23)

II) Electrons in Atoms (Chapter 13)
  A) Review Rutherford's Model
  B) Bohr's Model
  C) Quantum mechanical model (Schrodinger & Heisenburg)
  D) Atomic electron orbitals (s,p,d & f) and electron configurations
  E) Identifying valence electrons

III) Ionic Bonding (Chapter 15)
  A) Valence Electrons (read p413-414, Problems p 418#1,3)
  B) Octet Rule (read p414-418, Problems p418#2,4, 5,6)
  C) Ionic Bonding (read p419-421, p 421#7)
  D) Properties of Ionic Compounds (p422-425, Problems p 425#9-13)
  E) Properties of Metallic Bonds and Metals (p427-428, Problems p429#15,17;)

(SOL) Learning Objectives

1. (4e) Identify a process as endothermic or exothermic based on whether it absorbs or releases heat.
2. (5f)Memorize and use $q=mC\ \Delta T$ to solve specific heat capacity and calorimetry problems.
3. (5e) Calculate energy changes during phase changes using molar heat of fusion, molar heat of vaporization,
4. (5d) Identify freezing point, ΔH_{fusion}, $\Delta H_{vaporization}$, and boiling point on a heating curve of water.
5. (2f) Determine the # of valence electrons and electron configurations for anions and cations
6. (3d) Explain why ionic bonds form in terms of electron transfer and the octet rule
7. (3d) Explain why Hydrogen, Lithium, and Beryllium break the octet rule in ionic compounds
8. (3d) Predict which compounds will be ionic based on their position on the periodic table.
9. (2h) Define an electrolyte
10. (2h) Predict which compounds will be electrolytes
11. (2h) Illustrate what happens when an ionic compound dissolves in water.
12. (2h) Explain why metals conduct electricity
13. Identify the contributions of Bohr, de Broglie, Planck, Heisenberg and Shrodinger to the development of the modern atomic model.
14. Use the Pauli Exclusion Principle, the Aufbau Principle, and Hund's Rule to determine electron configurations.
15. Identify the shapes of the s, p and d orbitals and the number of electrons in each.
16. Provide the spdf orbital electron configuration of elements using the periodic table.

Chapter 11 Noteguide Part 1: Energy and Changes of State

I. ENERGY CHANGES

A. Definitions

Energy: capacity to do work (movement), or supply heat

Infrared

Potential Energy: stored energy (food, fuel, wood... any fuel)

Kinetic Energy: energy due to motion $KE = \frac{1}{2}mv^2$ ← "like speed"

"Therm...") Heat: A form of Ⓔnergy... transferred from HOT → COLD (temp = measure of heat)

Thermochemistry: study of heat changes due to chemical processes or phase changes

B. Exothermic and Endothermic Processes

Processes that absorb (endo) heat or release (exo) heat.

Definitions

Consider $NH_4Cl(s) \xrightarrow{water} \overbrace{NH_4^+(aq) + Cl^-(aq)}^{dissolving}$

ammonium chloride
\+

•NH_4Cl dissolved, and ∴ absorbed energy from H_2O

(endothermic)
(get colder)

T_1 of H_2O = 20°C

T_2 for H_2O = 10°C

"H" Enthalpy: ← synonym for "heat" @ constant pressure. (heat in a system.

System: The part of the Universe that you're studying

Surroundings: Everything else.

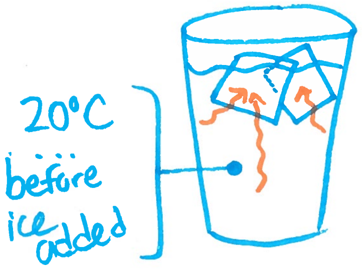
ΔH: "change in heat" or "change in enthalpy" (in the system)

Endothermic Process: Heat flows INTO system $\boxed{+\Delta H}$

Exothermic Process: Heat flows OUT of system. $\boxed{-\Delta H}$

$E=mc^2$

Another example of Endothermic and Exothermic processes
ice melting in glass of water:



System: $\boxed{\text{ice cubes}}$ (@ -5°C)

Surroundings: water (@20°C)

(heat from into ice)

- $H_2O(s) + HEAT \rightarrow H_2O(l)$
- $H_2O(s) \xrightarrow{\Delta} H_2O(l)$

Law of Conservation of Energy: Energy is neither created nor destroyed... only TRANSFERRED!

II. HEAT CAPACITY AND SPECIFIC HEAT CAPACITY

A. Definitions

Joule (J): SI unit of energy one kJ = 1000 J ($\times 1000$)

1 calorie: Amount of energy needed to raise the temp. of 1.0 g of H_2O by 1°C.

→ 1000 "science" calorie = 1 Calorie (food labels) and 1 calorie = 4.184 J

10/31

11/1

Specific Heat Capacity: The amount of heat (energy) needed to raise the temp of 1 g of something by 1°C

$\boxed{g/L}$

Symbol: C

Units: $\boxed{\frac{J}{g \cdot °C}}$ or $\frac{J}{g \cdot K}$

Equation:

q = heat (J or cal)

C = specific heat capacity

m = mass

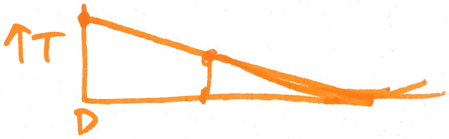
ΔT = change in temperature.

$q = mC\Delta T$

$T_i = 22°C$ $T_f = 12°C$

$\Delta T = T_f - T_i = 12°C - 22°C$

$= \boxed{-10°C}$



$q = mC\Delta T$

$\boxed{\Delta T = T_2 - T_1}$

$\Delta T = T_f - T_i$

B. Solving Specific Heat Capacity Problems

$$q = m\,C\,\Delta T$$

The equation has four variables: "q" is heat in Joules; "m" is mass in grams; "C" is specific heat capacity in J/(g•°C); "ΔT" is change in temperature in °C (the change in temperature is the final temperature minus the initial temperature, or $\Delta T = T_f - T_i$). This equation is only valid if the substance does not change phases. Identify the variables, then solve for the missing variable.

Sample Problems

1. A 500 g sample of iron changes from 22.0°C to 35.0°C. The specific heat of iron is known to be 0.46 J/(g•°C). How much heat was added?

m

T_i & T_f …∴ $\Delta T = \boxed{+13°C}$

$q = mC\Delta T$

$= (500)(0.46)(+13)$

$= 2{,}990\ J$

2. A 500. g sample of water changes from 22.0° to 35.0°C. The specific heat of water is known to be 4.18 J/(g•C) How much heat was added?

m

ΔT

$q = mc\Delta T$

$q = (500)(4.18)(+13)$

$q = 27{,}170\ J$

3. When 82 J of heat is added to a sample of aluminum, its temperature increased by 15.3°C. Given that the specific heat capacity of aluminum is 0.90 J/(g•°C), what is the mass of the sample?

ΔT

$q = \boxed{m}C\Delta T$

$m = \frac{q}{C\Delta T} = \frac{82}{(.90)(15.3)} = 5.95\ g$

$\Delta T = T_2 - T_1$

$\Delta T = T_f - T_i$

4. It takes 78.2 J to raise the temperature of 45.6 g lead by 13.3°C. Calculate the specific heat capacity of lead.

5. ~~Challenge:~~ A 142 g sample of silver at a temperature of 19.6°C absorbs 61.30 J of heat. What is the final temperature of the sample? [C_{Ag} = 0.24 J/(g•°C)] (*Ans = 21.4°C*)

m T_1 +q C

$q = mC\Delta T$

$\boxed{\Delta T = T_2 - T_1}$

$q = mC(T_2 - T_1)$

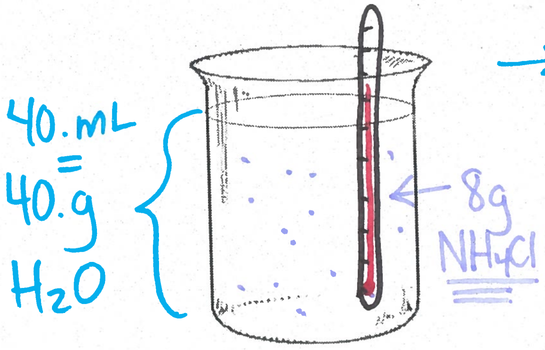
$\frac{q}{mC} = T_2 - T_1 \rightarrow \boxed{\frac{q}{mC} + T_1 = T_2}$

$T_1 = 21.4°C$

C. Calorimetry

Calorimetry: The msmt of heat energy change during a physical or chemical process. (food, fuel)

Calorimeter: Instrument used to measure calories.



$m_{H_2O} = 40.g$

$C_{H_2O} = 4.184 \frac{J}{g\cdot C}$

$\Delta T_{H_2O} = -10°C$

(J)

$q_{H_2O} = (40)(4.184)(-10)$

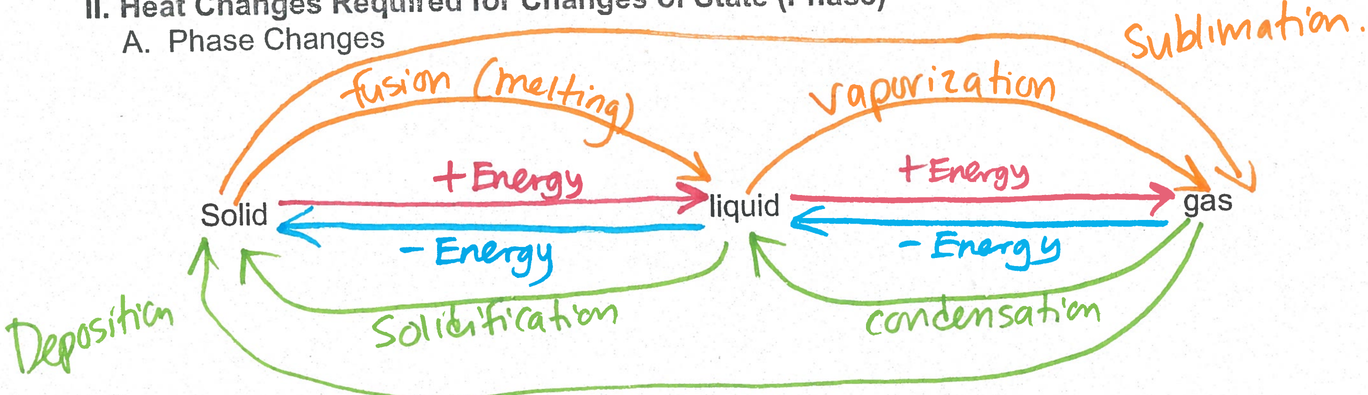
$q = -1673.6\ J$

$-q_{H_2O} = q_{NH_4Cl}$

$-(-1673.6) = \boxed{+1673.6\ J}$

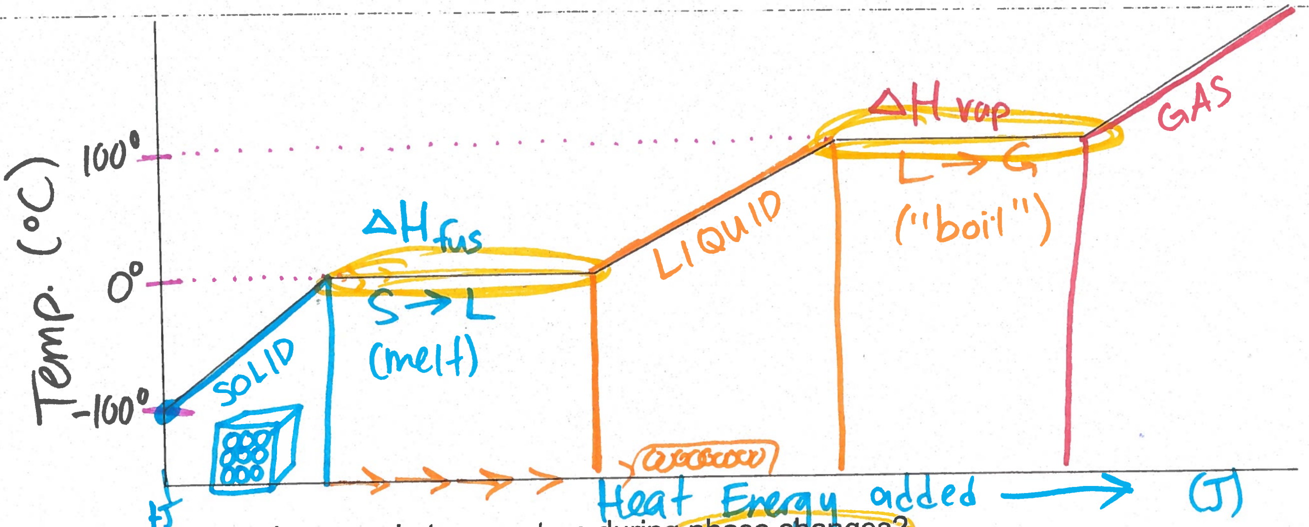
II. Heat Changes Required for Changes of State (Phase)

A. Phase Changes



Heat is absorbed or released when matter changes state.

Interpreting Heat Curve for water Relates T & Heat for Phase Changes



Does water increase in temperature during phase changes?

NO.

fancy word for melting

S → L

B. Molar Heats of Phase Changes

Molar heat of fusion: $\Delta H_{fusion} = \Delta H_{fus}$: Heat absorbed when 1 mole of something melts.

ENERGY amount: J/g or J/mol or cal/g or cal/mol

opp

L → S

Molar Heat of Solidification: $\Delta H_{solid.}$: Heat released when 1 mole of something solidifies

$$\boxed{\Delta H_{fusion} = - \Delta H_{solidification}}$$

Molar heat of vaporization: ΔH_{vap} : Amt of heat absorbed when 1 mole vapes.

opp.

Molar Heat of condensation: $\Delta H_{cond.}$: Amt of heat released when G → L

$$\boxed{\Delta H_{vap.} = - \Delta H_{cond.}}$$

Molar Heats apply to phase changes The units may include:

J/mol J/gram calories/gram

possibilities are, like, endless.

Solve the problems as unit analysis problems

C. Using Molar Heats in Calculations.

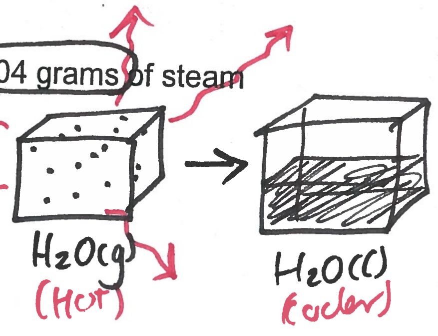
Example 1: How many grams of ice would be melted by adding 2.25 kJ of heat to an ice cube at 0°C? ΔH_{fusion} = 6.0 kJ/mol a conversion factor H_2O

$$\frac{2.25\ \cancel{kJ}}{1} \times \frac{1\ \cancel{mol\ ice}}{6.0\ \cancel{kJ}} \times \frac{18.0\ g\ ice}{1\ \cancel{mol\ ice}} = 6.75\ g\ H_2O(s)$$

Example 2: How many kilojoules of heat would be released when 36.04 grams of steam condenses to water at 100˚C?

$H_2O(g) \rightarrow H_2O(l)$ and $\Delta H_{cond.}$ = -40.7 kJ/mol

$$\frac{36.04\ \cancel{g\ H_2O(g)}}{1} \times \frac{1\ \cancel{mol\ H_2O}}{18.0\ \cancel{g\ H_2O}} \times \frac{-40.7\ kJ}{1\ \cancel{mol}}$$



Molar Heat Calculations Practice

molar heat is given in J/mol or J/g or cal/g, so use it as a conversion factor

1. How much heat is required to melt 500.9 grams of ice at O°C? The heat of fusion of water is 80.0 cal/g.

energy / amount

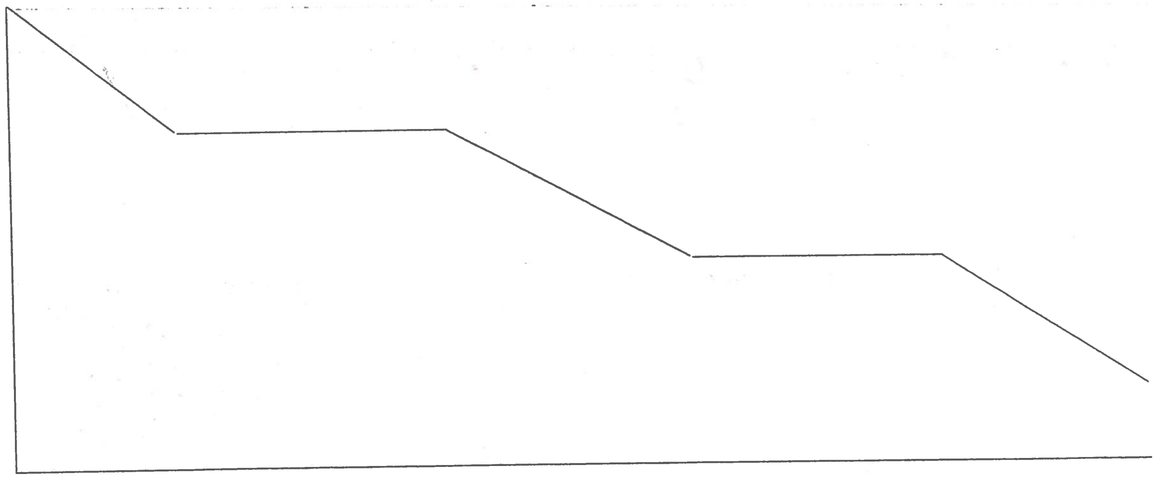
$$\frac{500.9\ \cancel{g}\ H_2O_{(s)}}{1} \times \frac{80.0\ cal\ H_2O_{(s)}}{1\ \cancel{g}\ H_2O_{(s)}}$$

2. How much heat is required to vaporize 13.1 grams of methane (CH_4) at it's boiling point, which has a heat of vaporization of 8.2 kJ/mol?

$$\frac{13.1\ \cancel{g}\ \cancel{CH_4}}{1} \times \frac{\cancel{mol}}{\cancel{g}} \times \boxed{\frac{kJ}{\cancel{mol}}}$$

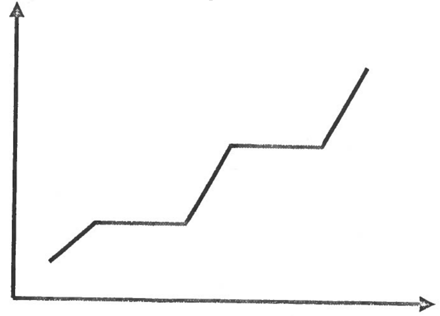
3. How many grams of neon must crystallize (solidify) at it's freezing point to release 560 J of heat, given that the neon's ΔH_{fusion} = 330 J/mol?

Interpreting a Cooling Curve for Water



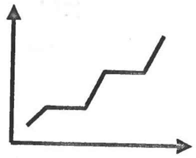
Mixed Molar Heats and Specific Heat Capacity Problems

Use the heating curve to tell which is which

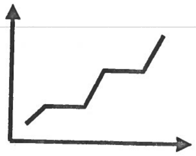


Specific heat capacity $H_2O(s)$ = 2.1 J/(g°C)
Specific heat capacity $H_2O(l)$ = 4.2 J/(g°C)
Heat of fusion H_2O = 6.0 kJ/mol
Heat of vaporization H_2O = 41 kJ/mol

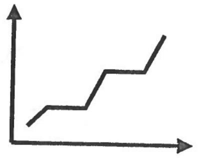
1. How much energy is needed to raise the temperature of 150 grams of ice from -20.0°C to -5.0°C? *(Ans = 4725 J)*



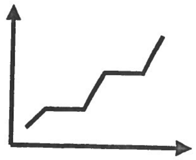
2. How much energy is needed to vaporize 52 grams of water at 100°C? *(Ans = 118 ≈120 kJ).*



3. How many grams of ice at 0°C would be melted by adding 820 kJ of heat. *(Ans = 2500 g ice)*

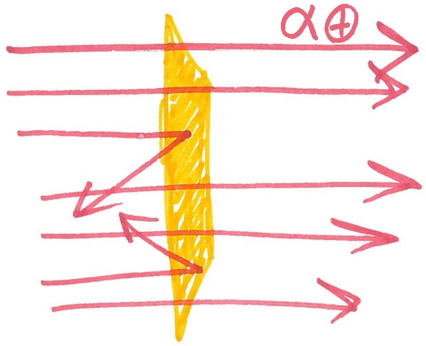
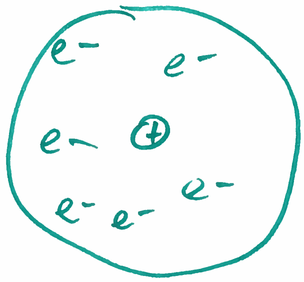


4. How much will the temperature of 850 grams of water increase if 16,000 Joules of heat is added? *(Ans = 4.5°C)*



Chapter 13: Electrons in Atoms

Review of Rutherford's Atomic Model(1911)

What	*How*	*Model*
Discovered the nucleus! Atom is mostly empty space. (tiny ⊕-ly charged nucleus)		

A. Important Terms

6
C
12.01

1. atomic number: number of protons-whole number shown on the periodic table 6
2. mass number: number of protons plus neutrons p+ ⊕ n° Nitrogen-15 7p+ 8n°
3. isotopes: elements with the same number of protons, but a different number of neutrons
4. atomic mass: weighted average of isotope masses. Listed on the periodic table.

B. Symbols for Isotopes neutral (p=e)

1. $^{13}_{6}C$ 6 protons, 6 electrons, 7 neutrons
2. ^{64}Cu 29 protons, 29 electrons, 35 neutrons
3. Pb−202 ____ protons, ____ electrons, ____ neutrons

OYO

C. Practice

ISOTOPE	ATOMIC #	# PROTONS	# NEUTRONS	MASS #
^{54}Fe				
	36		40	
		13	14	

How many electrons, neutrons and protons in Zinc-67?

How many neutrons are in F-19?

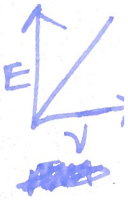
An Aside About Light and Energy

Light is fast. It travels at 3.0×10^8 m/s. (*distance* over *time* is speed, which is the magnitude of velocity). "c" is the constant that represents light's speed in a vacuum.

$E = mc^2$

or "f" Light **frequency** (ν, called "nu") × Light **wavelength** (λ, called "lambda") = c

Max Planck (1900) determines Said "Energy of light is directly proportional to its frequency" $E \propto \nu$ $E \propto f$



Equations:

Direct

Energy of light, using frequency:

$$E = \underset{\text{constant}}{\underset{\uparrow}{h}} f = h\nu$$

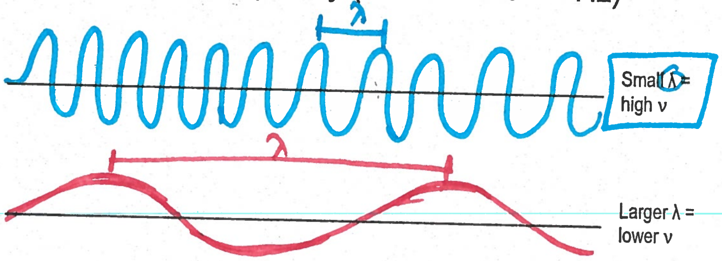
Inverse.

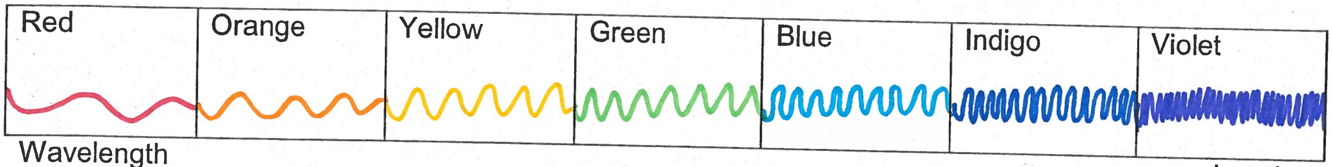
Energy of light, using wavelength:

$$E = h \frac{c}{\underset{\uparrow}{\lambda}}$$

$h = 6.626 \times 10^{-34}$ J·s

ν ("nu" not "v") = frequency (unit: $1/s \equiv s^{-1} \equiv$ Hz)





Wavelength
$= 7.0 \times 10^{-7}$ m
Low Energy

wavelength
4×10^{-7} m
High Energy

V-
P
D

Louis de Broglie (1924) determines All matter has wave-like properties. BIG particles = small waves.

SMALL particles = big waves

The de Broglie Equation

$$\lambda = \frac{h}{\boxed{m \cdot v}} = \frac{h}{\boxed{p}}$$

→ momentum

& Interpretation:

X

"h" = Planck's constant

6.63×10^{-34} J·s

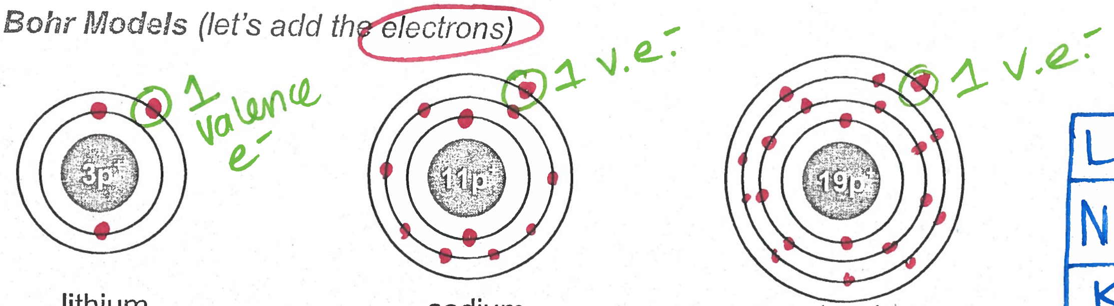
What did the spectrum tube demonstration show?

All elements have a unique spectrum... because they have unique electron confiruations.

Bohr's Model of the Atom (Powerpoint)

What	*How*	*Model*

A Few Bohr Models (let's add the electrons)



lithium sodium potassium

Li
Na
K

The number of electrons in the outer principal energy level (or valence shell) is the same within a group.

The number of principle energy levels is the same as the period #.

So why isn't this model good enough?

The Quantum Mechanical Model

A. Erwin Schrödinger used complex mathematics to calculate where electrons **probably** are around the atom. His mathematical models were revolutionary to physics.

Kinetic Energy + Potential Energy = E

Classical
Conservation of Energy
Newton's Laws

$\frac{1}{2}mv^2 + \frac{1}{2}kx^2 = E$ Harmonic oscillator example.

$F = ma = -kx$

Quantum
Conservation of Energy
Schrodinger Equation

$\frac{p^2}{2m} + \frac{1}{2}kx^2$ The energy becomes the Hamiltonian operator

$H\Psi = E\Psi$ ← Wavefunction

Energy "eigenvalue" for the system.

In making the transition to a wave equation, physical variables take the form of "operators".

$p \to \frac{\hbar}{i}\frac{\partial}{\partial x}$ $x \to x$

$H \to \frac{-\hbar^2}{2m}\frac{\partial^2}{\partial x^2} + \frac{1}{2}kx^2$ The form of the Hamiltonian operator for a quantum harmonic oscillator.

Both alive & dead.

B. Werner Heisenberg adds the **Heisenenberg Uncertainty Principle:**

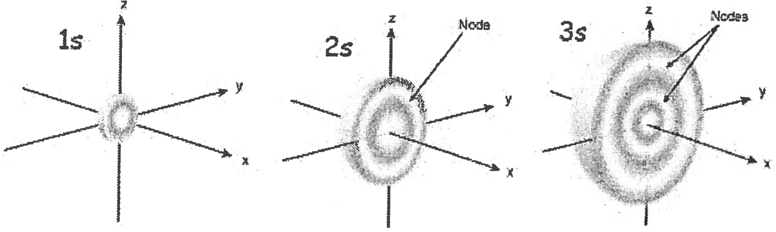
Can't know the position AND velocity of an electron
→ know only 1.

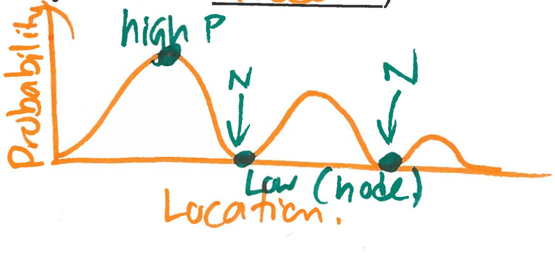
$\Delta x . \Delta p \geq \frac{h}{4\pi}$

$\Delta p = m\,\Delta v$

$\Delta x . m\Delta v \geq \frac{h}{4\pi}$

C. The areas where an electron can *probably* be found are called orbital cloud. (The areas where electrons will be unlikely are called nodes)





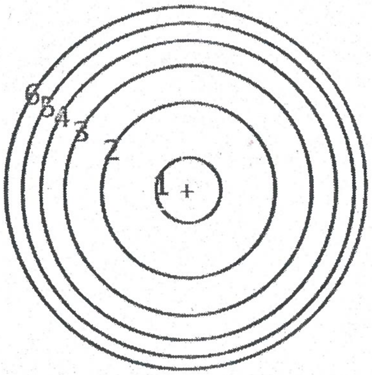
□

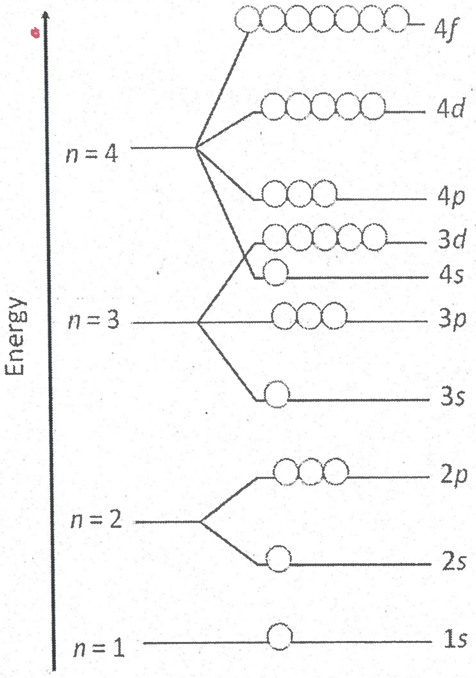
D. Each orbital has a specific shape and can hold up to 2e⁻ (spinning in ***opposite*** directions).

a. How is the spin of an electron noted in models? arrows

$\boxed{\uparrow\downarrow}$ or $\boxed{\uparrow\ \downarrow}$

E. Organization of Electrons





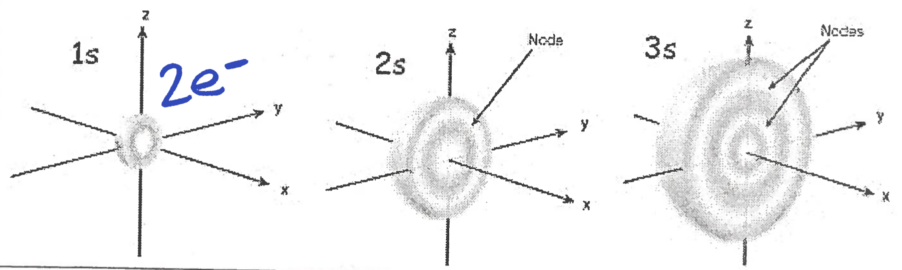
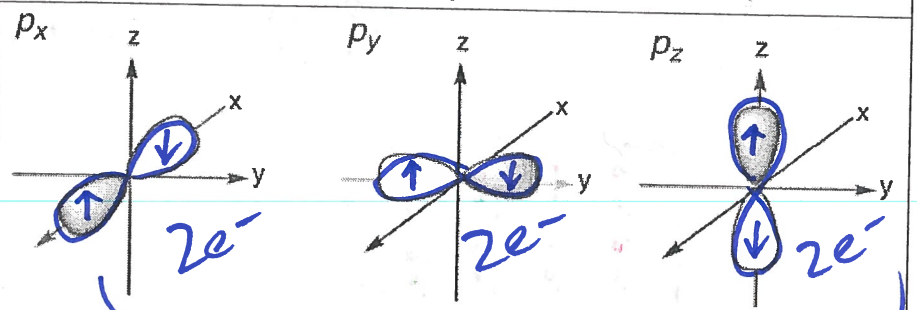
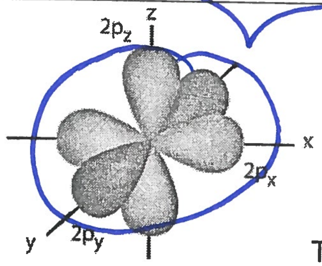
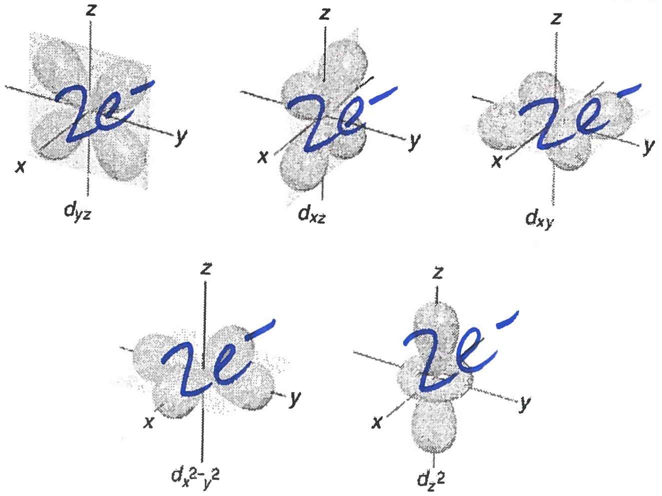
Description of Sublevels

1. "s" sublevels have 1 orbital – it is Sphere shaped

2. "p" sublevels have 3 orbitals – they are dumbell / hourglass shaped

3. "d" sublevels have 5 orbitals – 4 are Cloverleaf shaped, one "pacifier"
4. "f" sublevels have 7 orbitals – they are crazy. shaped

11/16
11/17

What orbital type is this? S How many electrons can go in this orbital? $2e^-$	 1s $2e^-$ 2s Node 3s Nodes z y x
How many electrons can fill each p- orbital? $2e^-$	 p_x p_y p_z z x y ↑ ↓ $2e^-$ The three p orbitals are aligned along perpendicular axes
How many electrons can fill this entire p-energy sublevel? $6e^-$	 $2p_z$ z x $2p_x$ y $2p_y$ The p-energy sublevel is made of all three 3D orientations (p_x, p_y, and p_z together)
How many electrons can fill each d-orbital? $2e^-$ How many electrons can fill this d-energy sublevel? $10e^-$	 z x y d_{yz} d_{xz} d_{xy} $d_{x^2-y^2}$ d_{z^2} $2e^-$

F. Filling in the Orbitals in Quantum Mechanics

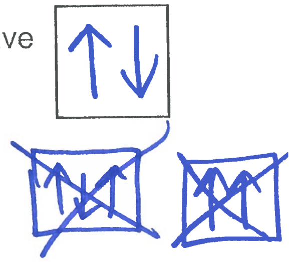
"to build up"

1. **Aufbau Principle:**

Electrons fill the lowest E-levels first.

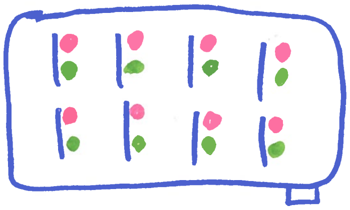
2. **Pauli Exclusion Principle** (PEP): ↑↓ or ↑↓

Only 2e⁻ electrons can be in each orbital (two per box or line!) and they must have opposite magnetic spins. (two different arrow directions!)



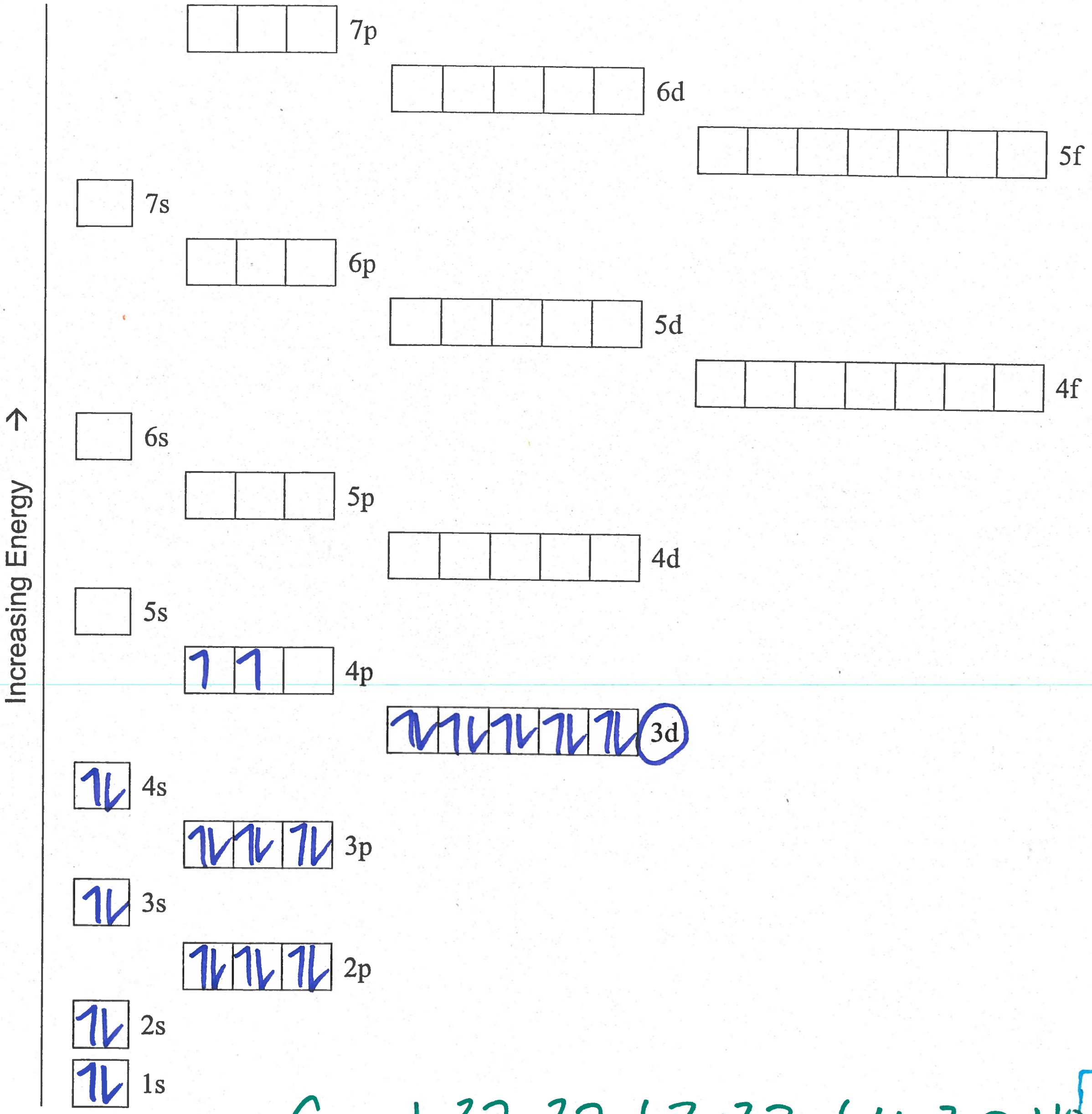
3. **Hund's Rule:**

When electrons occupy orbitals of equal energy, they fill in singly with aligned spins *before* they double up (space out if you can!) The bus seat analogy…



$32e^-$ in Germanium

Electron Configurations & the Aufbau Diagram



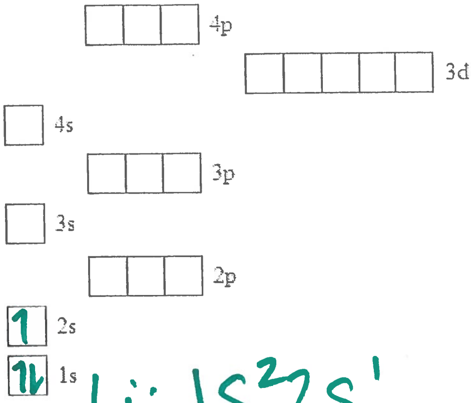
Ge: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} \boxed{4p^2}$

Rules to fill it in:

1. Electrons enter lowest energy first. (start with "1s") *[Aufbau Principle]*
2. An orbital can have at most 2 electrons with opposite spins. *[Pauli Exclusion Principle]*
3. When electrons are filling orbitals of equal energy, one electron enters each before they start to spin pair (double up). *[Hund's Rule]*

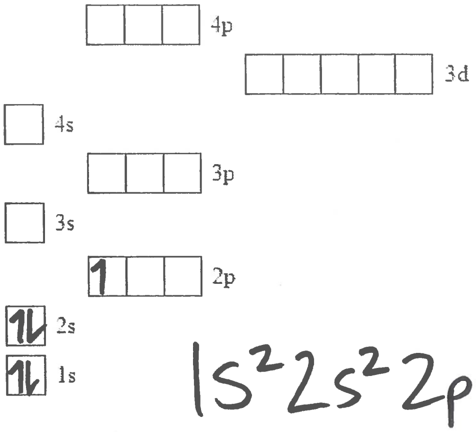
|||

Li #electrons = 3e⁻



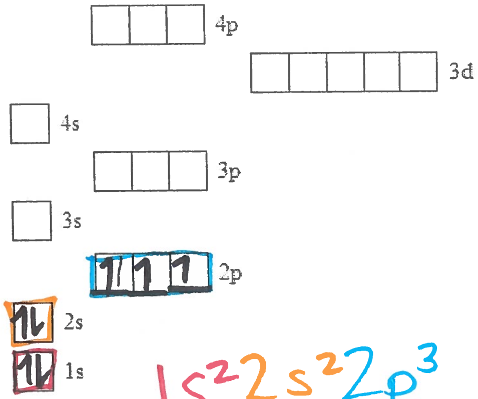
Li: $1s^2 2s^1$

B # electrons = 5e⁻



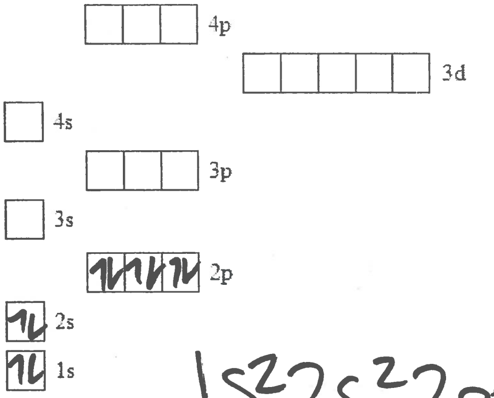
$1s^2 2s^2 2p^1$

N # electrons = 7e⁻



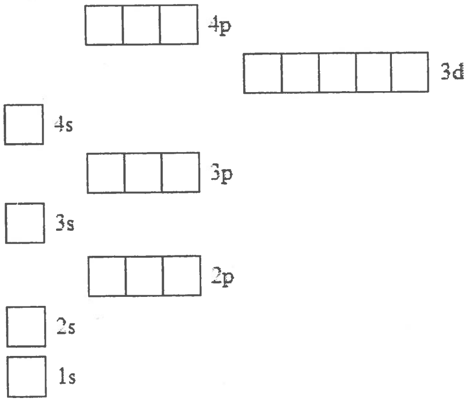
$1s^2 2s^2 2p^3$

Ne # electrons = 10e⁻

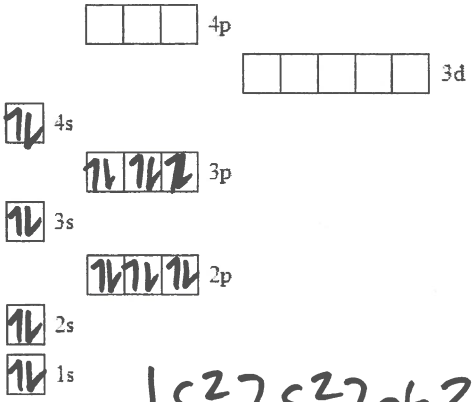


$1s^2 2s^2 2p^6$

P # electrons =



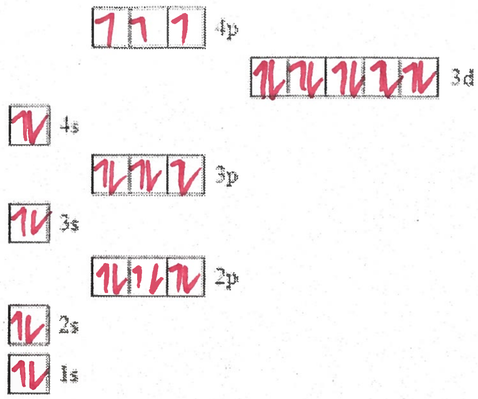
Ca #electrons = 20e⁻



$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

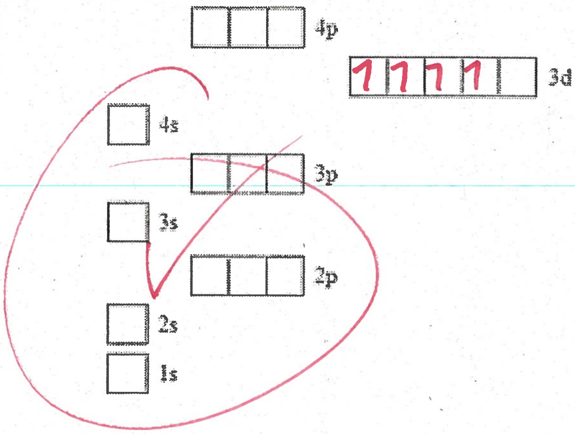
OYO

As # electrons = 33e⁻



OYO

Cr # electrons =



Complete the alternate form of the Aufbau diagrams below:

14e⁻ Si	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑								
	1s	2s	2p	2p	2p	3s	3p	3p	3p	4s	3d	3d	3d	3d	3d	
OYO Mg																
	1s	2s	2p	2p	2p	3s	3p	3p	3p	4s	3d	3d	3d	3d	3d	
OYO Fe																
	1s	2s	2p	2p	2p	3s	3p	3p	3p	4s	3d	3d	3d	3d	3d	

Indicate how many unpaired electrons each element has

Si: 2e⁻

Mg:______

Fe:______

Valence electron = outer e^-

$ns^x np^y$ and $x+y$ = # of valence e^-

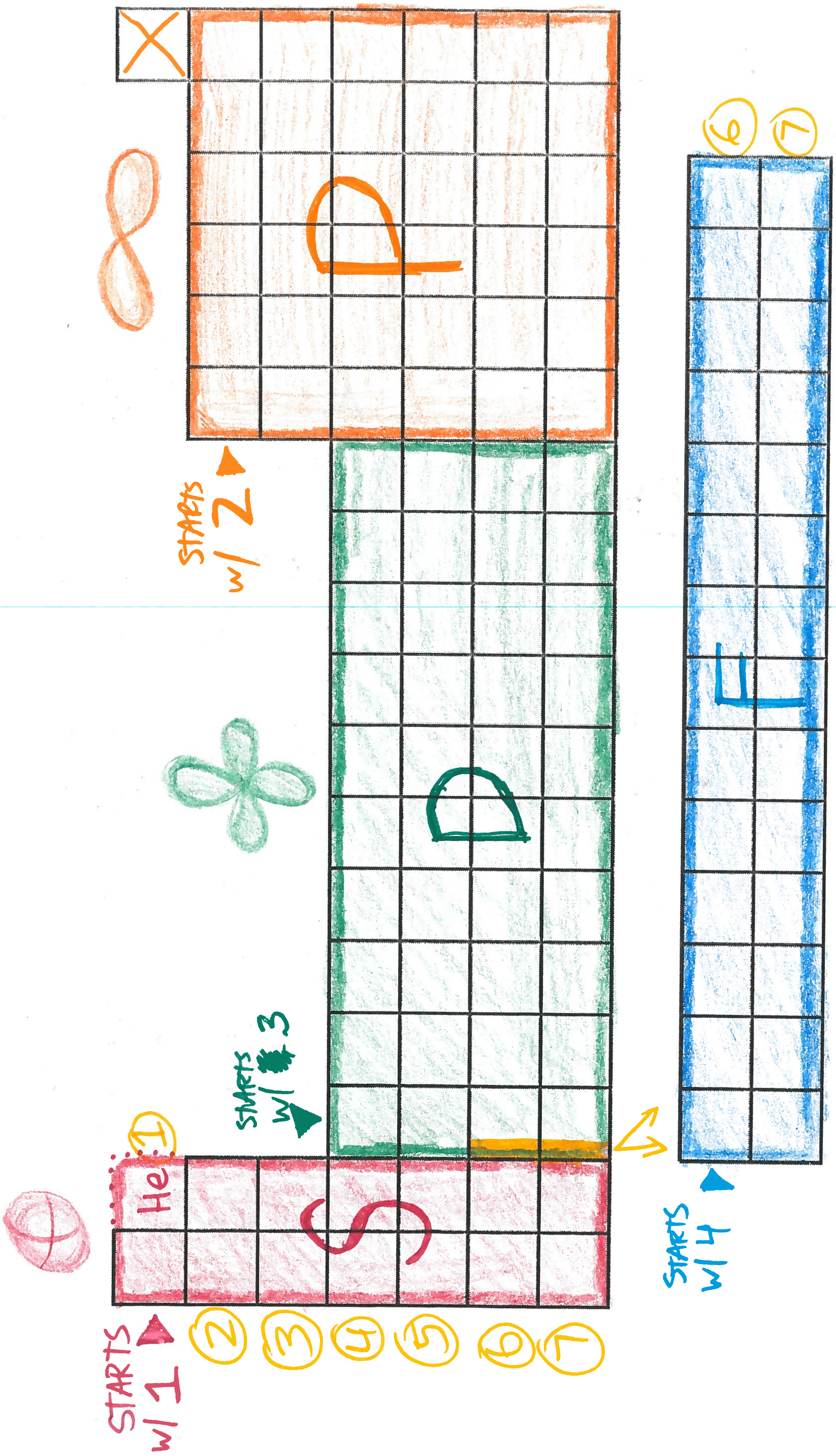
Chapter 13: Periodic Table and Electron Configuration

$:\ddot{\underset{\cdot\cdot}{F}}:$

1	H $1s^1$									He $1s^2$
group	1	2			13	14	15	16	17	18
valence e- config.	ns^1	ns^2			ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
2	Li	Be			B	C	N	O	F	Ne
3			Tran. Metals							
4			Sc	Zn						
5										
6					Tl					Rn
7	Fr	Ra	Lr	Uu b						
# valence electrons										

La	Yb
Ac	No

Representative elements: ______________________


S
He
1
2
3
4
5
6
7
STARTS w/ 1
P
STARTS w/ 2
D
STARTS w/ 3
F
STARTS w/ 4
6
7

Chapter 15: Ionic Bonds

I. **Valence** Electrons

A. definition: e- in outermost shell (higher energy!) These bond!

~~N~~

B. Lewis Dot Structures: show valence electrons as dots; the symbol Fe represents the core electrons (which is everything *but* the valence electrons).

~~S~~

Lewis Structures show bonded atoms as lines.

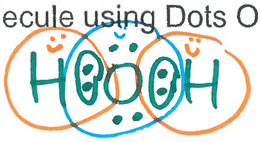
Lewis Dot Structure Example (single atom):

S

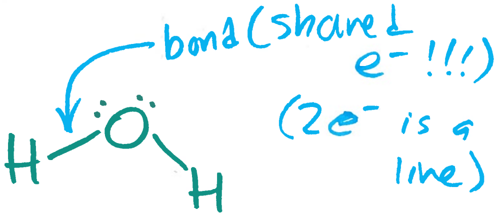
N (5v.e.) Cl (7ve) Ar (8ve) B

Lewis Structure of Molecule using Dots Only:

$1s^1 \rightarrow 1s^2$



Lewis Structure of Molecule using Lines for Bonds:



Group	1 Alkali	2 Alkaline	13 B↓	14 C↓	15 N↓	16 O↓	17 Halogen	18 N. Gases
Example	Na	Ca	B	Si	P	S	Cl	Kr

8 v.e. = ☺

II. **Octet** Rule

A. definition: atoms are generally-ish more stable w/ 8 v.e.- (try to look like "nearest NG")

B. A full valence shell is very stable (which is happy! ☺). Therefore, elements gain or lose electrons to reach a full octet.

8 →

$\boxed{ns^x np^y}$

- configuration example: Na = $1s^2 2s^2 2p^6 \boxed{3s^{1}}$ ← valence electron.

  Na^{+} = $1s^2 2s^2 2p^6$ ← "Argon"

:S·

- configuration example: S = $1s^2 2s^2 2p^6 3s^{2} 3p^{4}$

  S^{2-} = $1s^2 2s^2 2p^6 3s^{2} 3p^{6}$

  "Argon"

He: $1s^{②}$

C. ***Exceptions*** to Octet Rule in Ionic Compounds

Helium is happy with __2__ valence electrons, so we call this exception the __duet__ rule. Atoms with atomic numbers close to He (such as __H__, __Li__, and __Be__) will be happy with 2 electrons. They can't fit 8!

$1s^2 \cancel{2s^1}$ $1s^2 \cancel{2s^2}$

H^+: = $1s^0$ H^- = $\boxed{1s^2}$ — Li^+ = $\boxed{1s^2}$ — Be^{2+} = $\boxed{1s^2}$ — $\boxed{He}$

____

____

= equal

Terminology: **iso**- means __same__ or __one__ (think: isosceles triangle)
-electronic refers to the number of __# of e⁻__.

∴, what does isoelectronic mean? same # of e⁻

Concept Check:
- Are He and Ne isoelectronic with each other? __NO__ 2 ≠ 8 or 10
- Are O^{2-} and Ne isoelectronic with each other? __Yes__
- 10 ≠ 18 Are F^- and Cl^- isoelectronic with each other? __Nah__
- ⑱ ⑱ Are Cl^- and S^{2-} isoelectronic with each other? __Yes!!!__
- Which noble gas will iodine become isoelectronic to when an iodine atom is ionized? __Xenon__

I ~~54~~ 53 e⁻
vs
I^- 54e⁻

- Na+ will lose one electron, to become isoelectronic with __Neon__.
- Which alkaline metal (lose 3e⁻) is most likely to ionize to become isoelectronic with the noble gas Krypton? __Sr__
- Are Mg^{2+} and N^{3-} isoelectronic? __Yes__ (to Neon!)

↑ ↑

D. A couple of other octet rule **exceptions**:

Boron (B) *actually* prefers to have _________ valence electrons (and it's stable that way!), rather than 8 like many others.

Atoms from sulfur and beyond can sometimes have more than 9. This is called

_____________________.

III. Ionic Bonding

A. Question: *Where do anions get their extra electron(s) from anyway?*

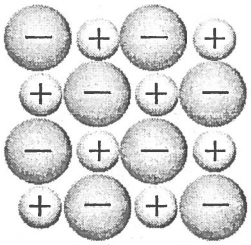
____

____

____

____

Examples: ($NaCl$, CaF_2, MgO, Li_3P, K_2S)



$\oplus \rightarrow \leftarrow \ominus$

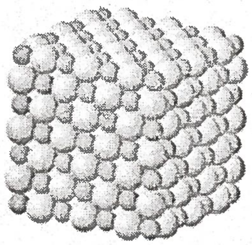
IV. Properties of Ionic Compounds

Ionic compounds are held together by strong electrostatic

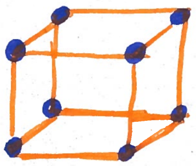
Electrostatic attraction: opp.-charged particles attract to each other!

$C_{12}H_{22}O_{11}$ sucrose

A. crystal structure: 3D orderly, repetitive crystal lattice arrangement of atoms...



NaCl



Na^+
K^+
Mg^{2+}
Ca^{2+}
Cl^-
I^-

B. electrolytes: Ionic compounds... that, when dissolved in H_2O, conduct electricity.

in body: • nerves & • muscles

C. high melting points

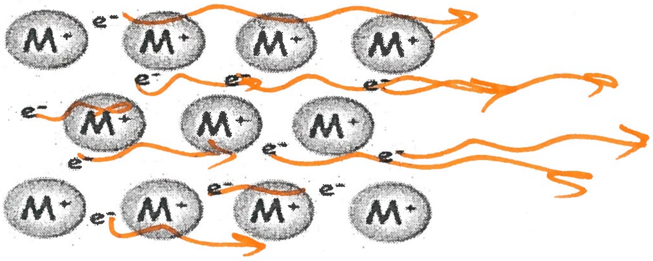
Ionic stuff... high MP

M + Nm

V. Metallic Bonds

A. caused by attraction of electrons (valence) for the positively charged nuclei in other atoms

B. metals are good conductors because of these free floating electrons.



metals

Ever notice how some metals, such as *steel*, *bronze*, and *brass* aren't on the periodic table???

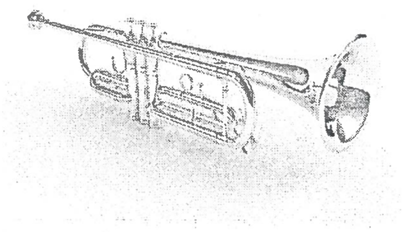
These are called ALLOYS. An **alloy** is a solid mixture of metals.

Two (or more) metals are melted, then mixed together while they're still liquid. After the hot liquid metal mixture cools, you have an alloy.

Brass is made of copper and zinc.
It's great for musical instruments due to how sound waves resonate (propagate) through the metal atoms!

Bronze is made of Copper, tin, and other metals.

Steel is made of iron, carbon, and other elements.




WHITE GOLD


YELLOW GOLD


PLATINUM

Jewelry… what is "white gold" and "rose gold"? Jewelry is often an alloy. White gold is an alloy of gold and another metal, like nickel or platinum. Metallic bonds keep it together, of course.

Investigation Questions:

Why is it not a good idea to have jewelry that is pure gold?

Pure gold is weak, brittle, malleable

What makes stainless steel special? And why doesn't it stain easily?

Coated with material that prevents oxidation

OYO Terms to Know:

Malleable ____________________

Conductor ____________________

Ductile ____________________

Brittle ____________________

Writing Ionic Formulas from Names Review

1. Identify the charge of the cation (see periodic table)
2. Use empty parentheses if you don't know the metal's charge immediately
3. Identify the charge of the anion
4. Identify the charge of the metal by canceling the anion's charges
5. Put the charge of the metal in the empty parenthesis. This is the *oxidation state* of the metal.

Magnesium carbonate

Calcium nitrate

Sodium phosphate

Tin (IV) chloride

Strontium Nitride

Copper (III) Sulfate

Naming Ionic Compounds Review

1. Name the cation
2. Does the cation name need a parentheses
3. Name the anion
4. Figure out the cations charge if needed

Li_2O ____________________

$CaCl_2$ ____________________

$Fe(NO_2)_3$ ____________________

Ba_3P_2 ____________________

$V(OH)_5$ ____________________

Cr_2O_3 ____________________

$Sr_3(PO_4)_2$ ____________________

$Cu(NO_3)_2$ ____________________

Chem ③ F'17

Chemistry Unit 4

Primary reference: ***Chemistry: Matter and Change*** **[Glencoe, 2017]**

Topic	Essential Knowledge	Study Support
Scientific Investigation 1.4 **SOL 1a,b,f**	Use chemicals and equipment safely. **Accuracy** is how close a measurement is to the true value. An accurate measurement has very little error. **Percent Error** = 100 x accepted value-exper. value /accepted value	
Atomic Structure and Periodic Relationships **2.3** **SOL 2d, 2g, 2i**	**Niels Bohr** proposed the planetary model of the atom with electrons located in distinct energy levels (orbits) around the nucleus. **Louis de Broglie** proposed that all particles have wavelengths. (including electrons). **Max Planck** proved that a photon's wavelength is proportional to its energy. **Schrodinger** calculated the theoretical shapes of electron orbitals (s,p,d,f). **Heisenberg** developed the uncertainty principle concerning an electron's location and velocity. Electrons are added one at a time to the lowest energy levels first (**Aufbau Principle**). An orbital holds a maximum of two electrons (**Pauli Exclusion Principle**). Electrons occupy equal-energy orbitals so that a maximum number of unpaired electrons results (**Hund's Rule**). Energy levels are designated 1–7. Orbitals are designated s, p, d, and f according to their shapes (sphere, dumbbell, 4-leaf clover.) The s, p, d, f orbitals relate to regions of the Periodic Table. **Valence electrons** occupy the highest principle energy level of an atom. All the elements in a group have the same number of valence electrons. An element's electron configuration determines the number of valence electrons. Example: Bromine's valence electron configuration is $4s^2 4p^5$ with 7 valence electrons. The outermost electrons in an atom are called **valence electrons**. The period (row) number on the periodic table corresponds to the outermost energy level occupied by the valence electrons in an element. Elements in the same group (column) on the periodic table have the same number of valence electrons **Lewis dot diagrams** show the valence electrons of an atom. The electrons (dots) are arranged around the element's symbol. **Metallic bonds** consist of the attraction of free-floating valence electrons for the positively charged metal ions.	**Ch 5:** Read pp. 136-145 on the electromagnetic spectrum; light, energy, and waves. Read pp. 144-148 on atomic emission spectra and electron energy levels. Read pp. 146-155 on development of the modern quantum mechanical model. Read p 154 on orbital shapes. Read pp 156-162 on electron configurations. & the aufbau principle Read pp. 161-162 on Lewis dot diagrams and valence electrons **Ch 7**: Read pp. 225-227 on metallic bonds
Nomenclature, Formulas, and Reactions **3.3** **SOL 3a, 3d, 3e**	Bonds form between atoms to achieve stability. Ionic compounds are formed by the attraction between positive and negative ions. Ions are formed by **electron transfer** from a metal to a non-metal (ionization). After electron transfer, both ions meet the **octet rule**. The octet rule is the tendency of an atom to take on the configuration of a noble gas.	**Ch 7:** Review/Read sections 7.1-7.3 (pp. 206-224) about ion formation, ionic compounds, and naming and writing formulas for ionic compounds.
Molar Relationships **4.4** **SOL 4d**	Dissolving is a physical change that involves heat. When an Ionic compound dissolves in water it breaks into the ions that make it up. This process is called **dissociation** and can be expressed by an equation. Example: $\mathbf{NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)}$ Ionic compounds that dissociate completely in water are strong electrolytes.	**Ch 9:** Read pp 299-300 (ions and dissociation) **Ch 14:** Read pp. 498-499 (electrolytes in solution)
Phases of Matter and Kinetic Molecular Theory **5.3** **SOL 5d,5e,5f**	**Specific Heat Capacity** (C) is a physical property of a substance. $\boxed{q = m \cdot C \cdot \Delta T}$ is used to calculate heat, mass, specific heat or temperature change respectively. Specific heat can be used to identify a substance. The equation is sometimes seen as $\mathbf{q = C \cdot m \cdot \Delta T}$; it means the exact same thing. **Atoms and molecules are in constant motion and they have more kinetic energy ("energy of motion") when they're hotter** (gas) versus cooler (liquid, then solid). Forces of attraction between molecules determine the physical changes of state. The intermolecular forces must be overcome in order for a substance to melt or boil. Phase changes that require heat (like melting or boiling) are **endothermic.** ΔH is positive for an **endothermic** change. This means heat goes in. **Molar heats of fusion and vaporization** can be used to calculate energy changes. Phase changes that give off heat (like freezing and condensing) are **exothermic.** ΔH is negative. This means heat is released. **Molar heats of solidification and condensation** can be used to calculate energy changes. Heating and cooling curves, known as **temperature line graphs,** show the energy changes that occur as a substance goes from a solid to a gas as temperature is changed.	**Ch 15:** Read and review pp 516-518 on heat energy. Read pp. 519-522 specific heat (capacity). Review pp. 525-528 on thermochemistry before reading pp. 529-533 on molar heats of phase changes.

Objectives for Unit 4

Chemistry: Matter and Change [Glencoe, 2017]

Topic Outline

I) Thermochemistry Part 1
  A) Types of Energy
  B) Exothermic and Endothermic processes
  C) Heat Capacity and Specific Heat
    1) Calculations using specific heat capacities
    2) Calorimetry
  D) Changes of State and Heat Changes
    1) Phase Changes and Interpreting Heating Curves
    2) Molar Heats of Fusion and Solidification
    3) Molar Heats of Vaporization and Condensation
II) Electrons in Atoms
  A) Review Rutherford's Model
  B) Bohr's Model
  C) Quantum mechanical model (Schrodinger & Heisenburg)
  D) Atomic electron orbitals (s,p,d & f) and electron configurations
  E) Identifying valence electrons
III) Ionic Bonding
  A) Valence Electrons
  B) Octet Rule
  C) Ionic Bonding
  D) Properties of Ionic Compounds
  E) Properties of Metallic Bonds and Metals

(SOL) Learning Objectives

1. (4e) Identify a process as endothermic or exothermic based on whether it absorbs or releases heat.
2. (5f) Memorize and use $q=mC\Delta T$ to solve specific heat capacity and calorimetry problems.
3. (5e) Calculate energy changes during phase changes using molar heat of fusion, molar heat of vaporization,
4. (5d) Identify freezing point, ΔH_{fusion}, $\Delta H_{vaporization}$, and boiling point on a heating curve of water.
5. (2f) Determine the # of valence electrons and electron configurations for anions and cations
6. (3d) Explain why ionic bonds form in terms of electron transfer and the octet rule
7. (3d) Explain why Hydrogen, Lithium, and Beryllium break the octet rule in ionic compounds
8. (3d) Predict which compounds will be ionic based on their position on the periodic table.
9. (2h) Define an electrolyte, and understand strong versus weak electrolytes.
10. (2h) Predict which compounds will be electrolytes
11. (2h) Illustrate what happens when an ionic compound dissolves in water.
12. (2h) Explain why metals conduct electricity
13. Identify the contributions of Bohr, de Broglie, Planck, Heisenberg and Shrodinger to the development of the modern atomic model.
14. Use the Pauli Exclusion Principle, the Aufbau Principle, and Hund's Rule to determine electron configurations.
15. Identify the shapes of the s, p and d orbitals and the number of electrons in each.
16. Provide the spdf orbital electron configuration of elements using the periodic table.

Thermochemistry, Part 1: Energy and Changes of State

I.ENERGY CHANGES

A. Definitions

Energy: Capacity to do $\overbrace{\text{work}}^{W=F\cdot d}$ or supply heat

Potential Energy: Stored energy (food; fuel; explosives)

$KE=\frac{1}{2}mv^2$ Kinetic Energy: "energy in/of motion" (moving particles)

Heat: form of energy (☆ JOULES ← unit (J))

↑ moving particles

→ Thermochemistry: study of energy changes in chem/physics (processes)

rxns

B. Exothermic and Endothermic Processes

☆ All combustion rxns are exothermic

Processes that absorb (endo) heat or release (exo) heat.

Definitions

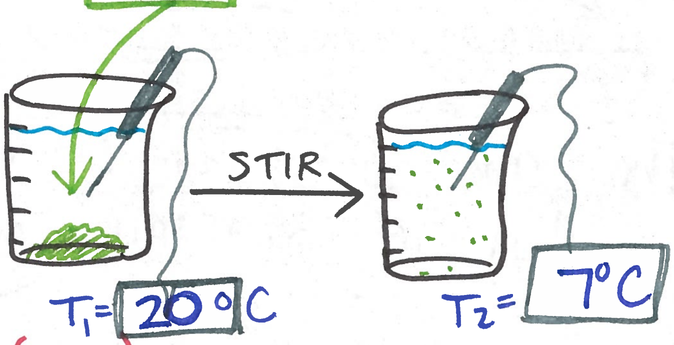
- Exothermic

A + B → C + heat

- endothermic

Heat + A + B → C

Consider $NH_4Cl(s) \xrightarrow{\text{water}} NH_4^+(aq) + Cl^-(aq)$



$\Delta T = -13°C$

Enthalpy: (ΔH) Heat in a system (@ constant P)

☆ System: Part of universe being "looked at"

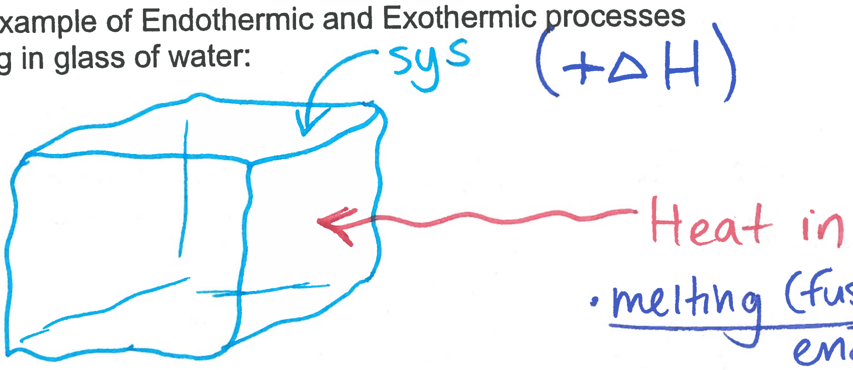
Surroundings: everything around system (else!!!)

ΔH: "change in" heat (enthalpy)

Endothermic Process: Heat flows INTO system $+\Delta H$

Exothermic Process: Heat flows OUT of system $-\Delta H$

Another example of Endothermic and Exothermic processes
ice melting in glass of water:



• melting (fusion) endothermic

Law of Conservation of Energy: Ⓔ cannot be destroyed or created… only transformed.

II. HEAT CAPACITY AND SPECIFIC HEAT CAPACITY

A. Definitions

☆ Joule (J): The SI unit for energy (J) one kJ = 1000 J

→ calorie: The amt energy needed to raise 1 g of H_2O by 1°C

1000 calories = 1 Calorie (food labels) = 1 kilocalorie (kcal)

also, 1 calorie = 4.184 J

Mm… A Snickers bar with 250 "Calories" (kilocalories) in food has 250,000 calories and 1,046,000 Joules.

☆ Specific Heat Capacity: The amt of heat (J) needed to raise 1 g of SOMETHING by 1°C

Symbol: C

Units: $\frac{J}{g \cdot °C}$

☆ Equation: $\boxed{q = mC\Delta T}$

q = → heat (enthalpy) (J)

C= specific heat capacity

m= mass (in g)

ΔT= "change in" temperature

$$q = mC\Delta T = Cm\Delta T$$

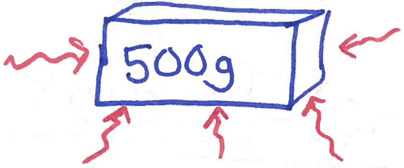
B. Solving Specific Heat Capacity Problems

$$q = m\,C\,\Delta T$$

The equation has four variables: "q" is heat in Joules; "m" is mass in grams; "C" is specific heat capacity in J/(g•°C); "ΔT" is change in temperature in °C (the change in temperature is the final temperature minus the initial temperature, or $\Delta T = T_f - T_i$). This equation is only valid if the substance does not change phases. Identify the variables, then solve for the missing variable.

Sample Problems

1. A 500 g sample of iron changes from 22.0°C to 35.0°C. The specific heat of iron is known to be 0.46 J/(g•°C). How much heat was added?



$m = 500g$

$\Delta T = +13.0°C$

$C = 0.46\ J/(g\cdot°C)$

$q = (500)(0.46)(13)$

$\therefore q = \boxed{2,990\ J}$

2. A 500. g sample of water changes from 22.0° to 35.0°C. The specific heat of water is known to be 4.18 J/(g•C) How much heat was added?

$m = 500g$

$\Delta T = +13.0°C$

$C = 4.18\ \left(\frac{J}{g\cdot°C}\right)$

$q = (500)(4.18)(+13)$

$\therefore q = \boxed{27,170\ J}$

OYO

3. When 82 J of heat is added to a sample of aluminum, its temperature increased by 15.3°C. Given that the specific heat capacity of aluminum is 0.90 J/(g•°C), what is the mass of the sample?

$m = ?$

$$\frac{q}{C\Delta T} = \frac{m\cancel{C\Delta T}}{\cancel{C\Delta T}}$$

$$m = \frac{q}{C\cdot\Delta T} = \frac{82}{(0.90)(15.3)} = \boxed{6.0g\ Al}$$

4. It takes 78.2 J to raise the temperature of 45.6 g lead by 13.3°C. Calculate the specific heat capacity of lead.

5. A 142 g sample of silver at a temperature of 19.6°C absorbs 61.30 J of heat. What is the final temperature of the sample? [C_{Ag} = 0.24 J/(g•°C)]

T_1

$T_2 = ?$

$q = mC\Delta T$ $\quad \Delta T = T_2 - T_1$

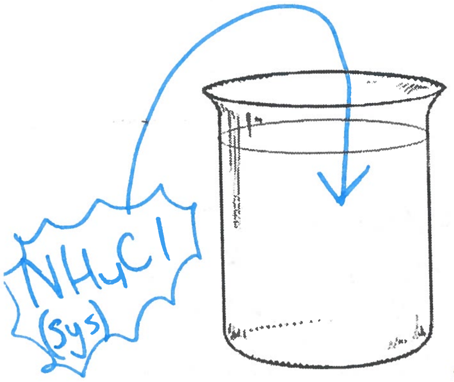
$q = mC(T_2 - T_1)$

C. Calorimetry

energy

Calorimetry: the process by which energy of a rxn is measured

Calorimeter: a device used to measure calories



$m_{H_2O} = \underline{40.\ g}$

$T_1 = 20°C\ ;\ T_2 = 9°C\ ;\ \therefore \Delta T = -11°C$

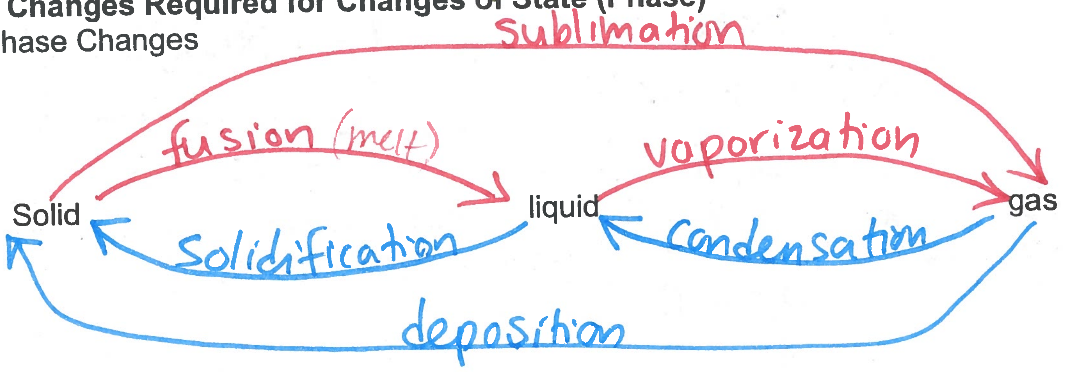
$C_{H_2O} = 4.184\ J/(g \cdot °C)$

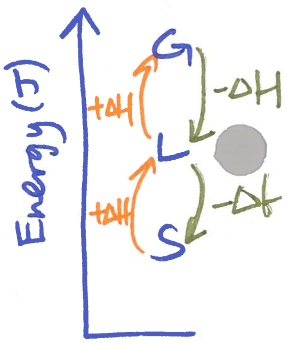
$\boxed{\boxed{q_{H_2O}} = \boxed{-q_{NH_4Cl}}} \rightarrow -(-1840) = \boxed{+1840\ J}$

$q_{H_2O} = (40\cancel{g})(4.184 \frac{J}{\cancel{g \cdot °C}})(-11\cancel{°C}) = \boxed{-1840\ J}$

II. Heat Changes Required for Changes of State (Phase)

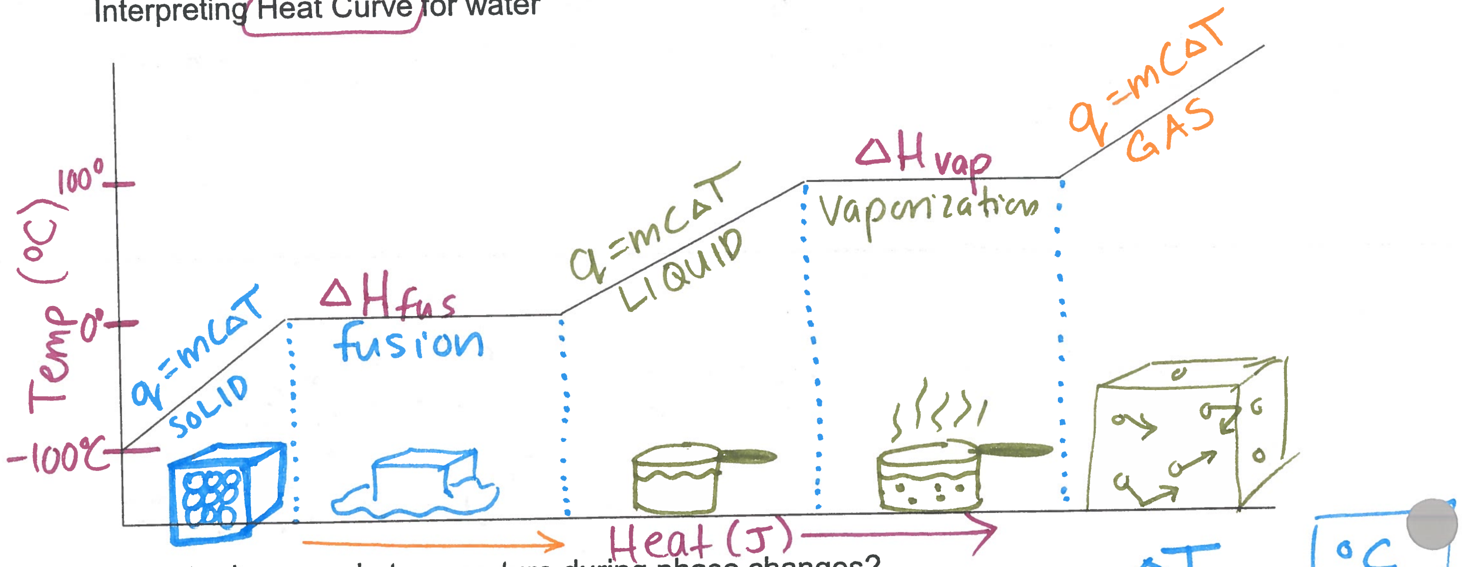
A. Phase Changes





Heat is absorbed or released when matter changes state.

Interpreting Heat Curve for water



Does water increase in temperature during phase changes?

No!!!

$\text{Slope} = \frac{\Delta T}{\Delta H} \rightarrow \boxed{\frac{°C}{J}}$

B. Molar Heats of Phase Changes

Molar heat of fusion: ΔH_{fus} : The amt of heat (J) needed to melt 1 mole of a substance

"Heat" over "amount"

J/g   J/mol   kJ/mol   cal/g

Molar Heat of Solidification: ΔH_{solid} : The amt of heat RELEASED when 1 mol freezes!

$\Delta H_{fusion} = -\Delta H_{solidification}$

Molar heat of vaporization: ΔH_{vap} : Heat needed for 1 mol of $\boxed{L \rightarrow G}$

Molar Heat of condensation: ΔH_{cond} : Heat released, 1 mol of $\boxed{G \rightarrow L}$

$\Delta H_{vap.} = -\Delta H_{cond.}$

Molar Heats apply to phase changes. The units may include:

J/mol   J/gram   calories/gram   kJ/mol

But they're ALWAYS HEAT *over* Amount*. Solve the problems as unit/dimensional analysis problems.*

C. Using Molar Heats in Calculations.

Example 1: How many grams of ice would be melted by adding 2.25 kJ of heat to an ice cube at 0°C? ΔH_{fusion} = 6.0 kJ/mol

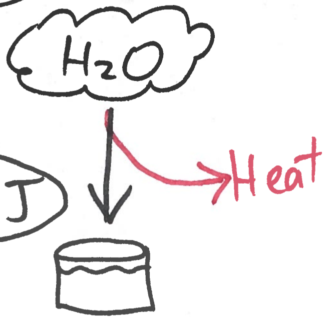
$$\frac{2.25\ \cancel{kJ}}{1} \times \frac{1\ \cancel{mol}}{6.0\ \cancel{kJ}} \times \frac{18.0\ g}{1\ \cancel{mol}} = \boxed{6.75\ g\ H_2O}$$

Example 2: How many kilojoules of heat would be released when 36.04 grams of steam condenses to water at 100˚C?

exo

$H_2O(g) \rightarrow H_2O(l)$ and $\Delta H_{cond.}$ = -40.7 kJ/mol

$$\frac{36.04\ \cancel{g}}{1} \times \frac{1\ \cancel{mol}}{18.0\ \cancel{g}} \times \frac{-40.7\ kJ}{1\ \cancel{mol}} = -81.5\ kJ$$



Molar Heat Calculations Practice

molar heat is given in J/mol or J/g or cal/g, so use it as a conversion factor

1. How much heat is required to melt 500.9 grams of ice at O°C? The heat of fusion of water is 80.0 cal/g.

$$\frac{500.9\ g}{1} \times \frac{80.0\ cal}{1\ g}$$

L→G

2. How much heat is required to vaporize 13.1 grams of methane (CH_4) at it's boiling point, which has a heat of vaporization of 8.2 kJ/mol?

$$\boxed{grams} \xrightarrow{MM} \boxed{mol} \xrightarrow{\Delta H_{vap}} \boxed{KJ}$$

$$\frac{13.1\ \cancel{g}\ CH_4}{1} \times \frac{1\ \cancel{mol}}{16.0\ \cancel{g}} \times \frac{8.2\ kJ}{1\ \cancel{mol}} = \boxed{}\ kJ$$

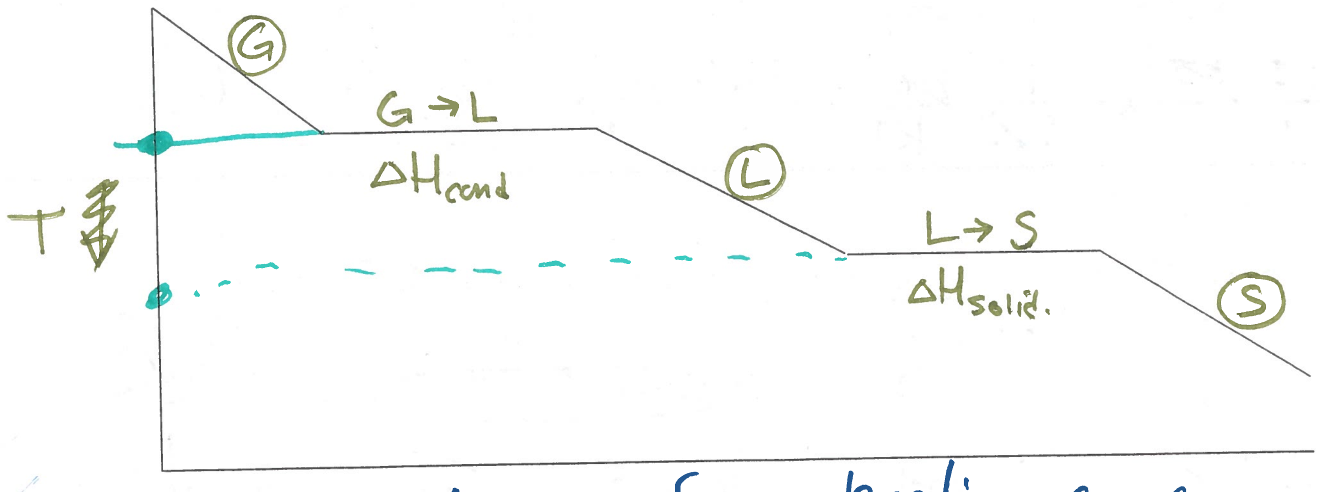
0/0

3. How many grams of neon must crystallize (solidify) at it's freezing point to release 560 J of heat, given that the neon's ΔH_{fusion} = 330 J/mol?

$$\boxed{J} \rightarrow \boxed{mol} \rightarrow \boxed{grams}$$

$$\frac{560\ J}{1} \times \frac{1\ mol}{330\ J} \times \frac{20.18\ g}{1\ mol} = \boxed{}\ g$$

Interpreting a Cooling Curve for Water

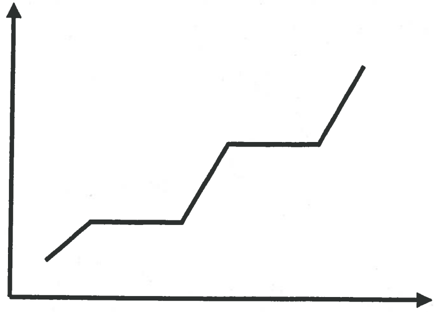


★ opp. of a heating curve

Heat →

Mixed Molar Heats and Specific Heat Capacity Problems

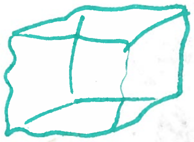
Use the heating curve to tell which is which



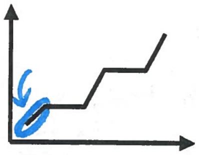
Specific heat capacity $H_2O(s)$ = 2.1 J/(g°C)
Specific heat capacity $H_2O(l)$ = 4.2 J/(g°C)
Heat of fusion H_2O = 6.0 kJ/mol
Heat of vaporization H_2O = 41 kJ/mol

1. How much energy is needed to raise the temperature of 150 grams of ice from -20.0°C to -5.0°C? *(Ans = 4725 J)*

$$q = mc\Delta T$$
$$= (150)(2.1)(+15)$$
$$= \boxed{4730\text{ J}}$$



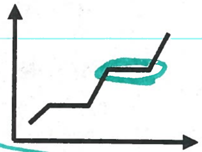
$\boxed{+4730}$ J



2. How much energy is needed to vaporize 52 grams of water at 100°C? *(Ans = 118 ≈120 kJ).*

$\boxed{L \to G}$

$$\frac{52\text{ g}}{1} \times \frac{1\text{ mol}}{18.0\text{ g}} \times \frac{41\text{ kJ}}{1\text{ mol}} = \boxed{120\text{ kJ}}$$

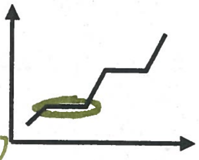


OYO

3. How many grams of ice at 0°C would be melted by adding 820 kJ of heat. *(Ans = 2500 g ice)*

$\boxed{\text{kJ}} \to \boxed{\text{moles}} \to \boxed{\text{grams}}$

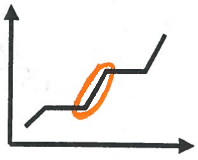
$$\frac{820\ \cancel{\text{kJ}}}{1} \times \frac{1\ \cancel{\text{mol}}}{6.0\ \cancel{\text{kJ}}} \times \frac{18.0\text{ g}}{1\ \cancel{\text{mol}}} = \boxed{2500\text{ g}}$$



OYO

4. How much will the temperature of 850 grams of water increase if 16,000 Joules of heat is added? *(Ans = 4.5°C)*

ΔT q



Chapter 13: Electrons in Atoms

Review of Rutherford's Atomic Model(1911)

What	***How***	***Model***

A. Important Terms

Z 1. atomic number: number of protons-whole number shown on the periodic table

M 2. mass number : number of protons plus neutrons

3. isotopes: elements with the same number of protons, but a different number of neutrons

4. atomic mass: weighted average of isotope masses. Listed on the periodic table.

B. Symbols for Isotopes

~~65 Cu 27~~

1. $^{13}_{6}$C 6 protons, 6 electrons, 7 neutrons

Cu-64 2. ^{64}Cu 29 protons, 29 electrons, 35 neutrons

3. Pb−202 82 protons, 82 electrons, 120 neutrons

C. Practice

53

Ⓐ Ⓑ Ⓒ

B=C-A

ISOTOPE	ATOMIC #	# PROTONS	# NEUTRONS	MASS #
^{54}Fe	26	26	28	54
^{76}Kr	36	36	40	76
$^{27}_{13}$Al	13	13	14	27

How many electrons, neutrons and protons in Zinc-67? OYO

How many neutrons are in F-19? OYO

~~C c~~ $E=mc^2$

An Aside About Light and Energy

Light is fast. It travels at 3.00 × 10⁸ m/s. (*distance* over *time* is *speed*, which is the magnitude of velocity). " C " is the constant that represents light's speed in a Vacuum.

ν = "nu"

Light **frequency** (f , called "_____") × **Light wavelength** (λ , called "lambda") = c

Max Planck (1900) determines Light color is related to the energy & frequency (rel. to λ) of the EM wave

$\frac{1}{x} = x^{-1}$

~~5~~

55

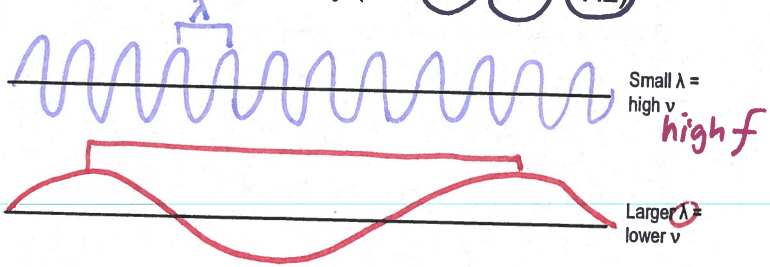
Period

Equations:

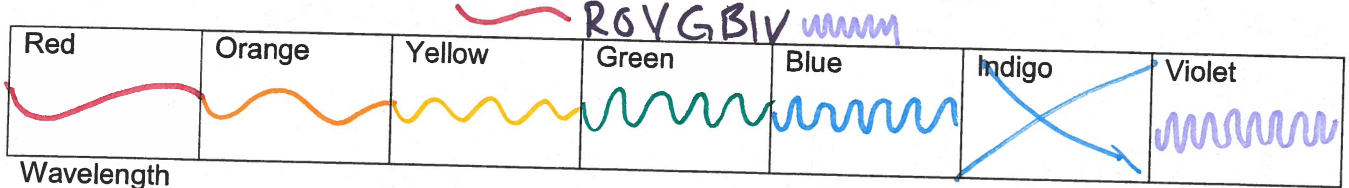
Energy of light, using frequency: $E = hf = h\nu$
Energy of light, using wavelength: $E = h\frac{c}{\lambda}$

$h = 6.626 \times 10^{-34}$ J·s

ν ("nu" not "v") = frequency (unit: 1/s = s^{-1} = Hz)



ROYGBIV



Wavelength = 7.0 x 10^{-7} m
Low Energy

wavelength
4 x 10^{-7} m
High Energy

11/8 →

Louis de Broglie (1924) determines ____________________

The de Broglie Equation
$\lambda = \frac{h}{mv}$

& Interpretation:

he spectrum tube demonstration show?

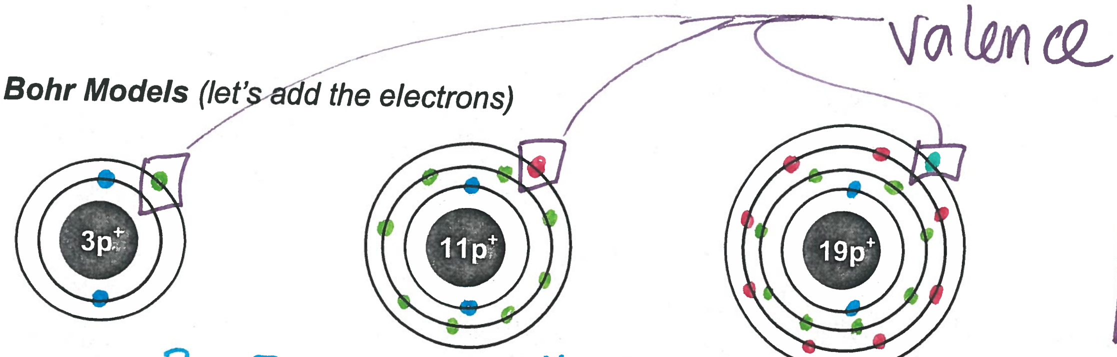
Bohr's Model of the Atom (Powerpoint)

What	***How***	***Model***

Group 1

Li
Na
K

A Few Bohr Models *(let's add the electrons)*



lithium 3e⁻ sodium 11e⁻ potassium 19e⁻

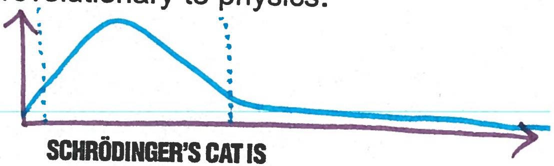
The number of electrons in the outer principal energy level (or valence shell) is the same within a group.

The number of principle energy levels is the same as the period #.

So why isn't this model good enough?

The Quantum Mechanical Model

A. Erwin Schrödinger used complex math to calculate where electrons **probably** are around the atom. His mathematical models were revolutionary to physics.



Kinetic Energy + Potential Energy = E

Classical
Conservation of Energy
Newton's Laws

$\frac{1}{2}mv^2 + \frac{1}{2}kx^2 = E$ Harmonic oscillator example.

$F = ma = -kx$

Quantum
Conservation of Energy
Schrodinger Equation

$\frac{p^2}{2m} + \frac{1}{2}kx^2$ The energy becomes the Hamiltonian operator

$H\Psi = E\Psi$ ← Wavefunction

Energy "eigenvalue" for the system.

In making the transition to a wave equation, physical variables take the form of "operators".

$p \rightarrow \frac{\hbar}{i}\frac{\partial}{\partial x}$ $x \rightarrow x$

$H \rightarrow \frac{-\hbar^2}{2m}\frac{\partial^2}{\partial x^2} + \frac{1}{2}kx^2$ The form of the Hamiltonian operator for a quantum harmonic oscillator.

position

B. Werner Heisenberg adds the **Heisenenberg Uncertainty Principle:**

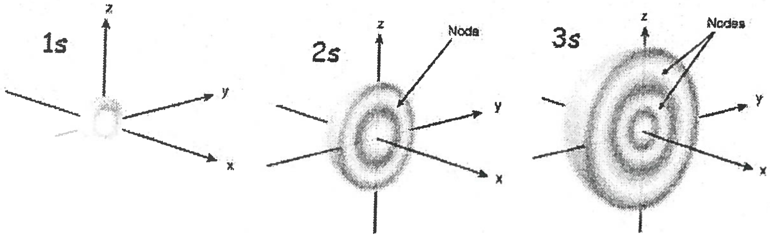
can't know LOCATION and
the momentum (velocity) ...
only 1

$\Delta x \Delta p \geq \frac{h}{4\pi}$

$\Delta p = m\Delta v$

$\Delta x m\Delta v \geq \frac{h}{4\pi}$

C. The areas where an electron can *probably* be found are called orbital. (The areas where electrons will be unlikely are called node)

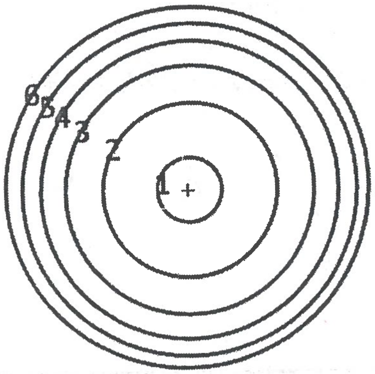


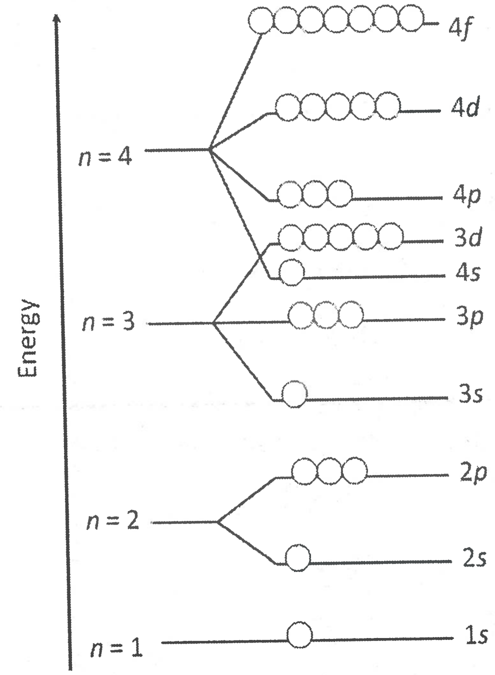
D. Each orbital has a specific shape and can hold up to 2 e^- (spinning in ***opposite*** directions).

a. How is the spin of an electron noted in models? arrows $\boxed{\upharpoonleft\downharpoonright}$

$\upharpoonleft\downharpoonright \equiv \uparrow\downarrow$

E. Organization of Electrons





Description of Sublevels

1. “s” sublevels have 1 orbital – it is Sphere shaped

2. “p” sublevels have 3 orbitals – they are dumbell shaped

3. “d” sublevels have 5 orbitals – 4 are Cloverleaf shaped, one “pacifier”

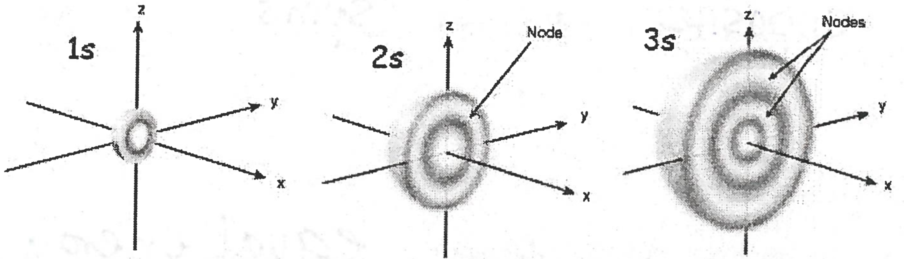
4. “f” sublevels have 7 orbitals – they are CRAZY shaped

What orbital type is this?

S

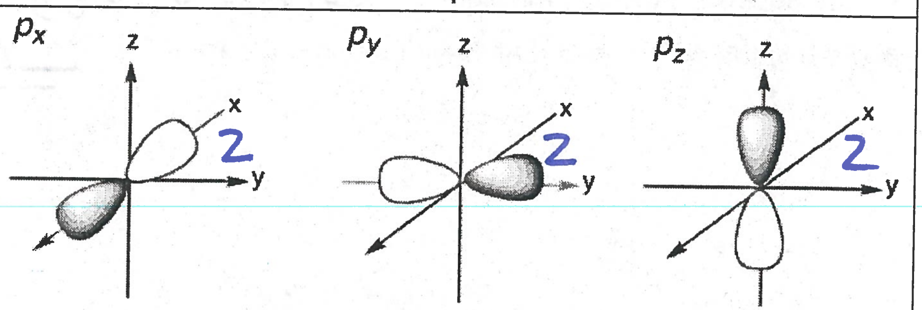
How many electrons can go in this orbital?

2



How many electrons can fill each p- orbital?

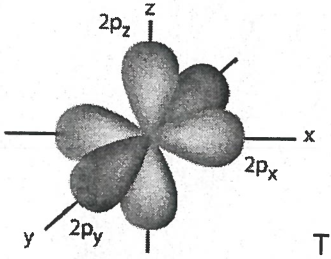
2



The three p orbitals are aligned along perpendicular axes

How many electrons can fill this entire p-energy sublevel?

6



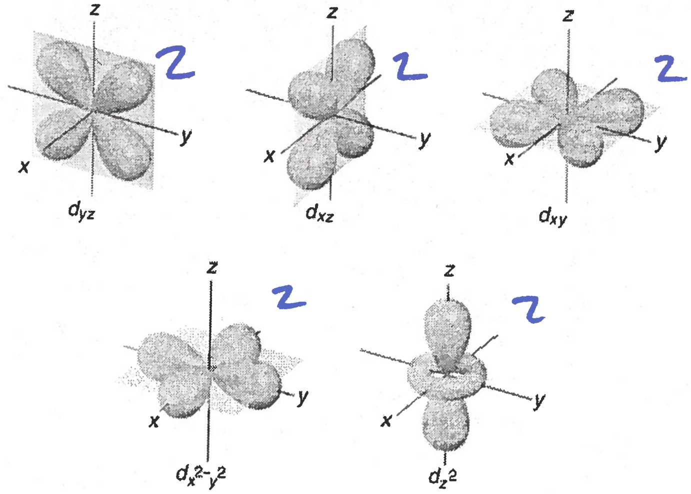
The p-energy sublevel is made of all three 3D orientations (p_x, p_y, and p_z together)

How many electrons can fill each d-orbital?

2

How many electrons can fill this d-energy sublevel?

10



F. Filling in the Orbitals in Quantum Mechanics

1. **Aufbau Principle**:

Electrons fill the lowest energy first first.

2. **Pauli Exclusion Principle** (PEP):

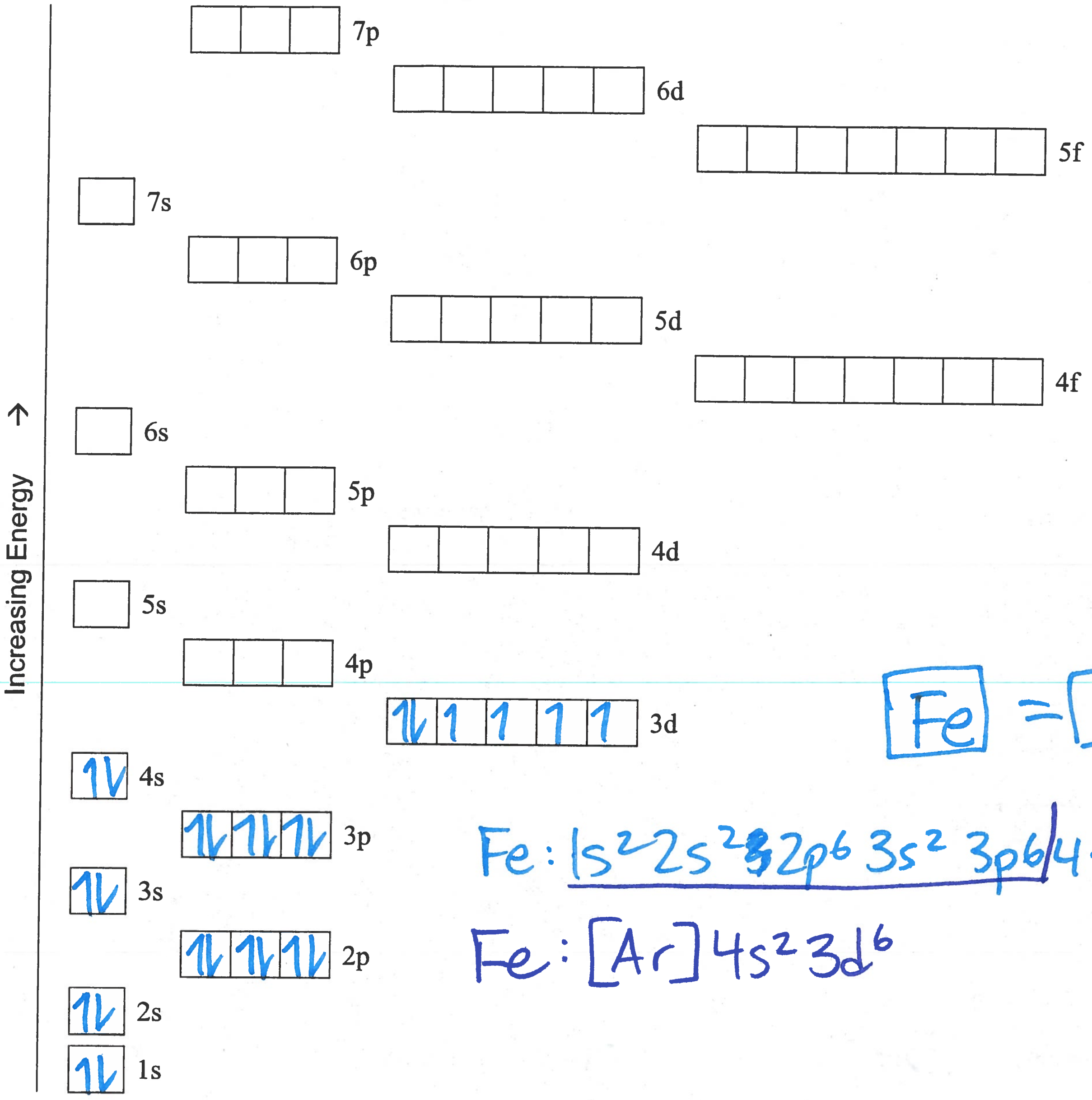
Only 2 electrons can be in each orbital (two per box or line!) and they must have opposite magnetic spins. (two different arrow directions!)

$\boxed{\uparrow\downarrow}$

3. **Hund's Rule**:

When electrons occupy orbitals of equal energy, they fill in singly with aligned spins *before* they double up (space out if you can!) The bus seat analogy…

Electron Configurations & the Aufbau Diagram



$\boxed{Fe} = \boxed{26} e^-$

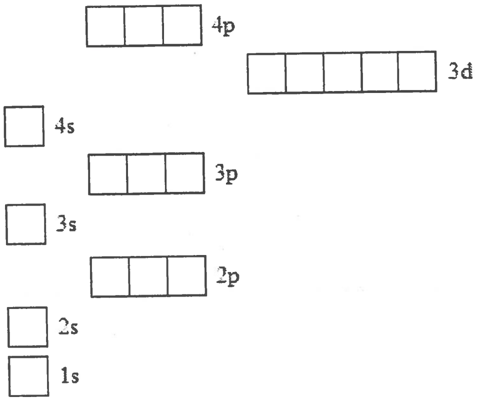
Fe: $\underline{1s^2 2s^2 2p^6 3s^2 3p^6} | 4s^2 3d^6$

Fe: $[Ar] 4s^2 3d^6$

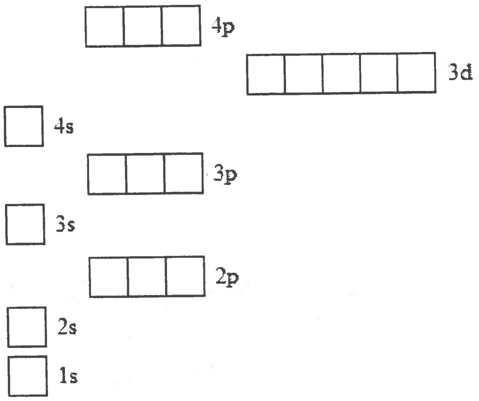
Rules to fill it in:

1. Electrons enter lowest energy first. (start with "1s") *[Aufbau Principle]*
2. An orbital can have at most 2 electrons with opposite spins. *[Pauli Exclusion Principle]*
3. When electrons are filling orbitals of equal energy, one electron enters each before they start to spin pair (double up). *[Hund's Rule]*

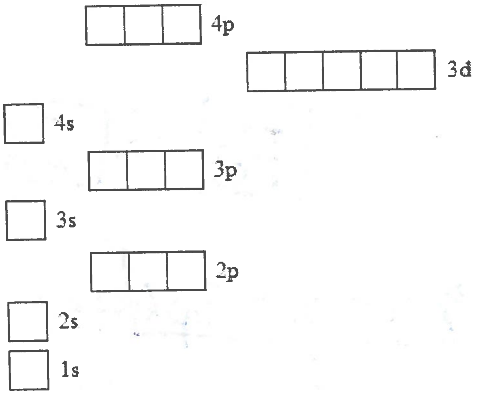
Li #electrons =



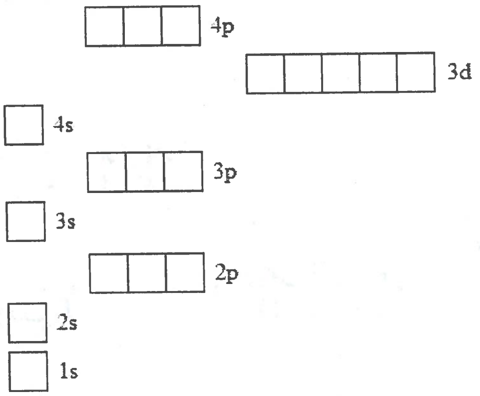
B # electrons =



N # electrons =

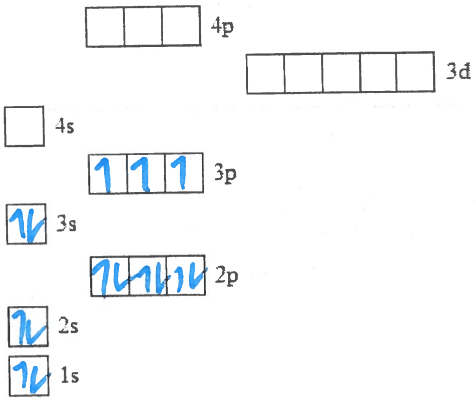


Ne # electrons =

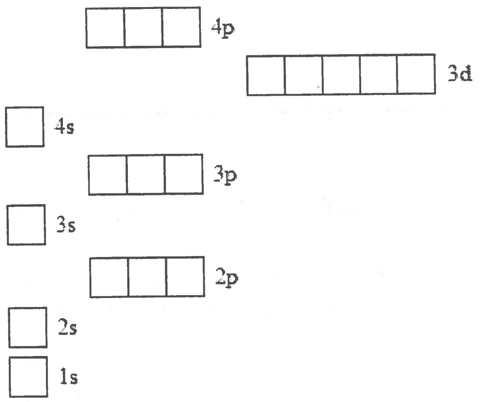


OYO

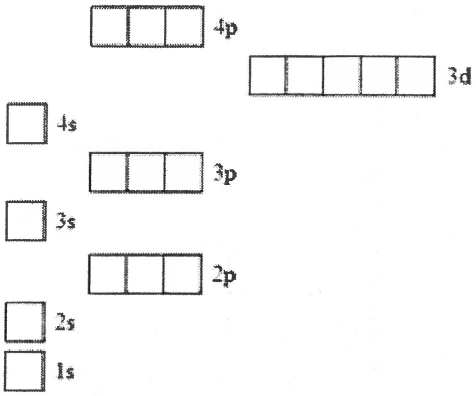
P # electrons = 15e⁻



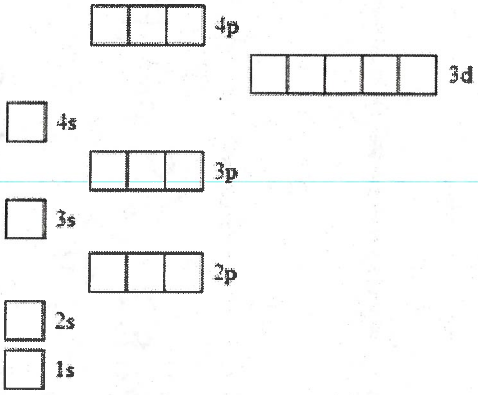
OYO Ca #electrons =



As # electrons =



Cr # electrons =



Complete the alternate form of the Aufbau diagrams below:

Si ___ 1s ___ 2s ___ 2p ___ 2p ___ 2p ___ 3s ___ 3p ___ 3p ___ 3p ___ 4s ___ 3d ___ 3d ___ 3d ___ 3d ___ 3d

✓ Mg ↑↓ 1s ↑↓ 2s ↑↓ 2p ↑↓ 2p ↑↓ 2p ↑↓ 3s ___ 3p ___ 3p ___ 3p ___ 4s ___ 3d ___ 3d ___ 3d ___ 3d ___ 3d

Fe ___ 1s ___ 2s ___ 2p ___ 2p ___ 2p ___ 3s ___ 3p ___ 3p ___ 3p ___ 4s ___ 3d ___ 3d ___ 3d ___ 3d ___ 3d

Indicate how many unpaired electrons each element has

Si: ______ Mg:______ Fe:_____

valence e⁻: outside e⁻... have highest energy! ★They Do Bonding!

Chapter 13: Periodic Table and Electron Configuration

$ns^{x}np^{y}$ ... $\boxed{x+y=\#v.e.^-}$

①	H																	He
group	1	2											13	14	15	16	17	18
valence e- config.	ns^1	ns^2											ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
②	Li	Be											B	C	N	O	F	Ne
period → ③																		
④			Sc									Zn						
⑤																		
⑥													Tl					Rn
⑦	Fr	Ra	Lr									Uub						
# valence electrons																		

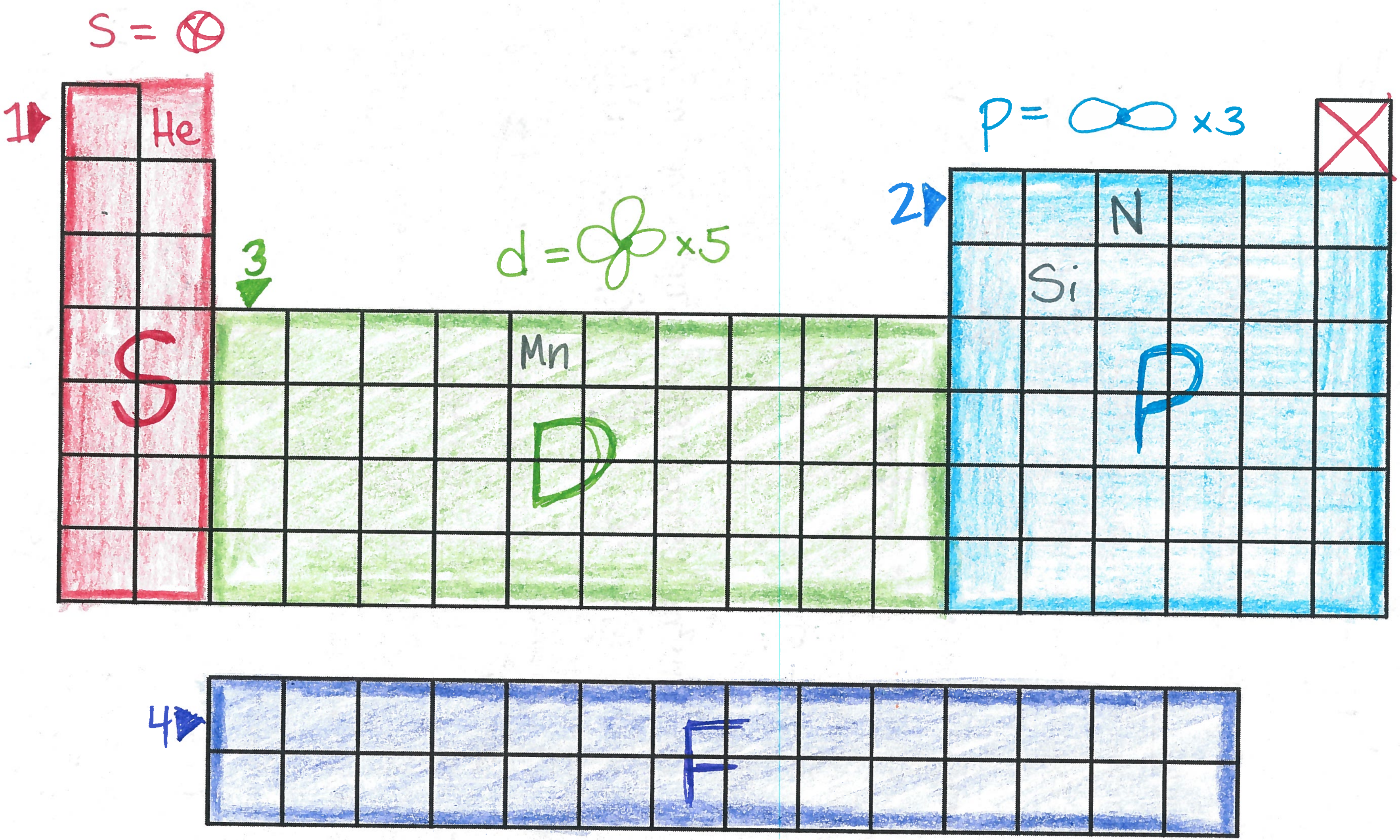
La													Yb
Ac													No

Representative elements: "TALL" columns (S and P)

S: [Ne] $3s^{②}3p^{④}$ 2 + 4 = $\boxed{6\ ve}$

★How many ve are in Al?

$1s^22s^22$[illegible] $\boxed{3s^23p^1}$



$P: \underline{1s^2 2s^2 2p^6}\boxed{3s^2 3p^3}$ $\underbrace{ns^{\square} np^{\square}}$

$P: \underline{[Ne]}3s^2 3p^3$

$P^{3-}: [Ne] 3s^2 3p^6 \Rightarrow [Ar]$

Chapter 15: Ionic Bonds

I. **Valence** Electrons

A. definition: high-energy, "outermost" e^-

B. Lewis Dot Structures: show valence electrons as dots; the symbol represents the core electrons (which is everything *but* the valence electrons).

$:\overset{..}{\underset{..}{X}}:$

Lewis Structures show bonded atoms as lines ← 2 ve⁻

sum $ns^{\square} np^{\square}$

Lewis Dot Structure Example (single atom): F : $1s^2 2s^{②} 2p^{⑤}$ = $\boxed{7ve^-}$

$\cdot\overset{..}{\underset{..}{F}}:$

Lewis Structure of Molecule using Dots Only: methane $\overset{④}{C}\overset{1+1+1+1}{H_4}$ = 8 ve⁻

Lewis Structure of Molecule using Lines for Bonds: $\overset{1+1+6}{H_2O}$ = 8ve⁻

$H-\overset{..}{\underset{..}{O}}-H$

He:

~~Al:~~ [illegible] ~~Ca:~~

Group	1	2	13	14	15	16	17	18
	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
Example	$K\cdot$	$\cdot Ca\cdot$	$\dot{Al}\cdot$	$\cdot\dot{\underset{.}{C}}\cdot$	$\cdot\ddot{\underset{.}{P}}\cdot$	$\cdot\ddot{\underset{..}{O}}\cdot$	$\cdot\ddot{\underset{..}{F}}:$	$:\ddot{\underset{..}{Ne}}:$

II. **Octet** Rule ☺

A. definition: "MOST" elements "want~~s~~" 8 v.e.⁻ (most stable)

H: $1s^1$

$\boxed{\uparrow}$ 1s

He: $1s^2$

$\boxed{\uparrow\downarrow}$ 1s

B. A full valence shell is very stable (which is happy! ☺). Therefore, elements gain or lose electrons to reach a full octet.

$ns^{\square}np^{\square}$

1 v.e.⁻

- configuration example: Na = $1s^2 2s^2 2p^6 \boxed{3s^1}$

$Na\cdot$

Na^{+} = $1s^2 \boxed{2s^2 2p^6} 3s^0$ 8.ve.⁻

~~H_2 vs He (2) (4)~~

- configuration example: S = $1s^2 2s^2 2p^6 \boxed{3s^2 3p^4}$ 6v.e.⁻

$:\overset{.\circ}{\underset{.\circ}{S}}:$

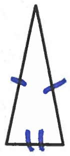
S^{2-} = $1s^2 2s^2 2p^6 \boxed{3s^2 3p^6}$ 8ve⁻

C. ***Exceptions*** to Octet Rule in Ionic Compounds

Helium is happy with 2 valence electrons, so we call this exception the duet rule. Atoms with atomic numbers close to He (such as H, Li, and Be) will be happy with 2 electrons. They can't fit 8!

H^+: = $1s^0$ H^- = $1s^2$ Li^+ = $1s^22s^0$ Be^{2+} = $1s^22s^0$

Terminology: **iso-** means Same or equal (think: isosceles triangle)
-electronic refers to the number of electrons.



∴, what does isoelectronic mean?

Concept Check:

- Are He and Ne isoelectronic with each other? No!
  $1s^2$ | $1s^22s^22p^6$
- Are O^{2-} and Ne isoelectronic with each other? yes!
  $1s^22s^22p^6$ | $1s^22s^22p6$
- Are F$^-$ and Cl$^-$ isoelectronic with each other? No!
  $1s^22s^22p^6$ | $1s^22s^22p^63s^23p^6$
- Are Cl$^-$ and S^{2-} isoelectronic with each other? yes!
  ⑱ = 17+1 | 16+2 = ⑱
- Which noble gas will iodine become isoelectronic to when an iodine atom is ionized?
  Xenon ← 54e- 53 ~~e~~ e-
- 11e- Na+ will lose one electron, to become isoelectronic with Neon (10 e-)
- Which alkaline metal is most likely to ionize to become isoelectronic with the noble gas Krypton? Sr
- Are Mg^{2+} and N^{3-} isoelectronic? Yes!

$\underbrace{1s^22s^22p^6}_{10e^-}\boxed{\cancel{3s^0}}$ | $\underbrace{1s^22s^22p^6}_{10e^-}$

D. A couple of other octet rule **exceptions**:

Boron (B) *actually* prefers to have 6 valence electrons (and it's stable that way!), rather than 8 like many others.

octet

Atoms from sulfur and beyond can sometimes have more than 9. This is called hyperValency.

III. Ionic Bonding

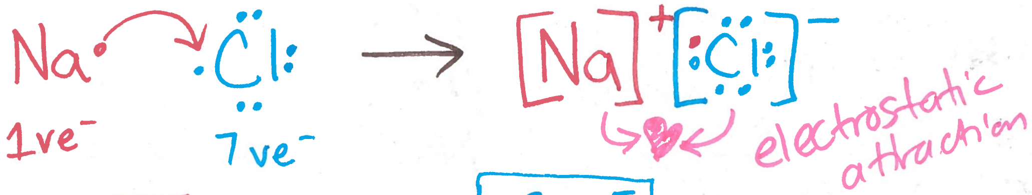
A. Question: *Where do anions get their extra electron(s) from anyway?*

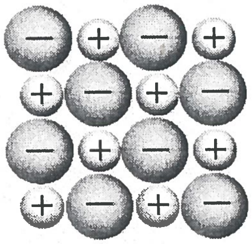
they get 'em from cations

↑ lose e^-

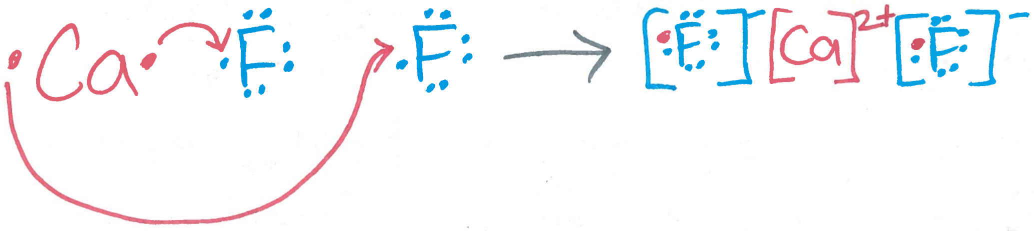
electron transfer

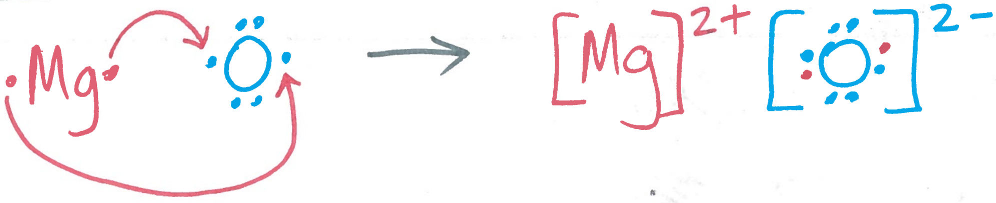
Examples: ($NaCl$, CaF_2, MgO, Li_3P, K_2S)



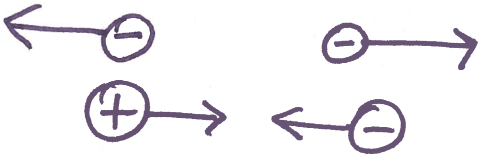


$1s^2 2s^2 2p^6 \boxed{3s^1}$ $1s^2 2s^2 2p^6 \boxed{3s^2 3p^5}$





Coulomb's Law

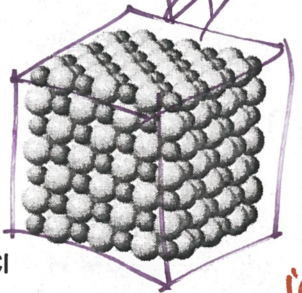


IV. Properties of Ionic Compounds

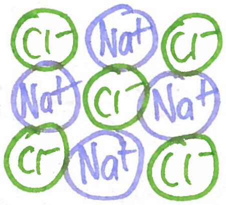
Ionic compounds are held together by electrostatic attraction

Electrostatic attraction: "opp. charges attract"

A. crystal structure: 3D pattern of repeating ⊕ & ⊖ charges



NaCl



B. electrolytes: ionic compounds that — when dissolved — conduct an electrical current

in H_2O

$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$

$\boxed{CaF_2}(s) \xrightarrow{H_2O} Ca^{2+}(aq) + 2F^-(aq)$

C. high melting points

↳ high MP... high AF

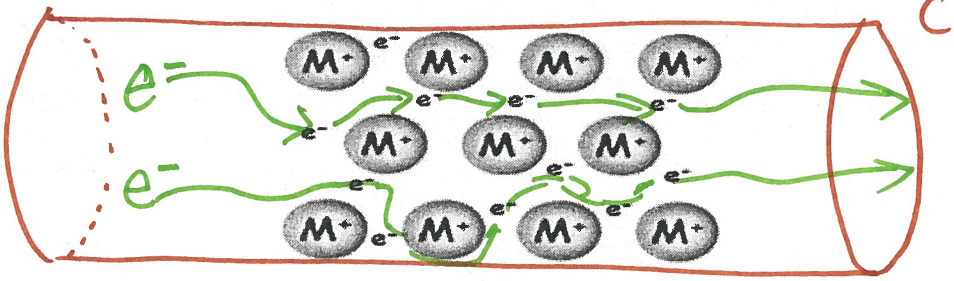
✩ ionic bonds are STRONG!!

V. Metallic Bonds

A. caused by attraction of free-floating electrons for the positively charged ~~[illegible]~~ nucleus

B. metals are good conductors because of these free floating electrons.

Cu(s)

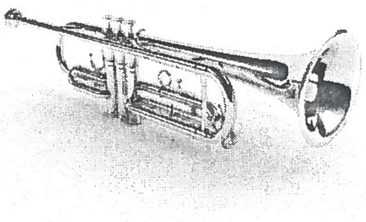


Ever notice how some metals, such as steel, bronze, and brass aren't on the periodic table???

These are called ALLOYS. An **alloy** is a solid mixture of metals.

Two (or more) metals are melted, then mixed together while they're still molten. After the hot liquid metal mixture cools, you have an alloy.

Brass is made of copper and zinc.
It's great for musical instruments due to how sound waves resonate (propagate) through the metal atoms!



Bronze is made of copper, tin, and other metals.

Steel is made of Iron, carbon, and other elements.

$4Fe + 3O_2 \rightarrow 2Fe_2O_3$

WHITE GOLD

YELLOW GOLD

PLATINUM

Jewelry… what is "white gold" and "rose gold"? Jewelry is often an alloy. White gold is an alloy of gold and another metal, like nickel or platinum. metallic bonds keep it together, of course.

24K

3 1

75mg Au + 25mg Cu

Investigation Questions:

Why is it not a good idea to have jewelry that is pure gold?

- Gold is malleable, brittle, easily dinged… durability

What makes stainless steel special? And why doesn't it stain easily?

- coating on surface that prevents oxidation…

OYO

OYO Terms to Know:

Malleable ______

Conductor ______

Ductile ______

Brittle ______

Phoenix

Writing Ionic Formulas from Names Review

1. Identify the charge of the cation (see periodic table)
2. Use empty parentheses if you don't know the metal's charge immediately
3. Identify the charge of the anion
4. Identify the charge of the metal by canceling the anion's charges
5. Put the charge of the metal in the empty parenthesis. This is the *oxidation state* of the metal.

← CAMEL

Magnesium carbonate

Mg^{2+} $CO_3^{2-} \rightarrow \boxed{MgCO_3}$

→ Sodium phosphate

Na^{+} PO_4^{3-} $\boxed{Na_3PO_4}$

Strontium Nitride

Calcium nitrate

Ca^{2+} NO_3^{-}

$Ca(NO_3)_2$

Tin (IV) chloride

Sn^{4+} $Cl^{-} \rightarrow SnCl_4$

Copper (III) Sulfate

Naming Ionic Compounds Review

1. Name the cation
2. Does the cation name need a parentheses
3. Name the anion
4. Figure out the cations charge if needed

Li_2O ____________________

$CaCl_2$ ____________________

→ $\overset{+++}{Fe}(\overset{---}{NO_2})_3$ iron(III) nitrite

Ba_3P_2 ____________________

→ $\overset{+5}{V}(\overset{-5}{OH})_5$ vanadium(V) hydroxide

Cr_2O_3 ____________________

$Sr_3(PO_4)_2$ ____________________

$Cu(NO_3)_2$ ____________________

Block 2
F'16

Chemistry Unit 4

Primary reference: *CHEMISTRY*, Addison-Wesley

Topic	Essential Knowledge	Study Support
Scientific Investigation **1.4** **SOL 1a,b,f**	Use chemicals and equipment safely. **Accuracy** is how close a measurement is to the true value. An accurate measurement has very little error. Percent Error = 100 x $\lvert$accepted value-exper. value$\rvert$ /accepted value	
Atomic Structure and Periodic Relationships **2.3** **SOL 2d, 2g, 2i**	**Niels Bohr** proposed the planetary model of the atom with electrons located in distinct energy levels (orbits) around the nucleus. **Louis de Broglie** proposed that all particles have wavelengths. (including electrons). **Max Planck** proved that a photon's wavelength is proportional to its energy. **Schrodinger** calculated the theoretical shapes of electron orbitals (s,p,d,f). **Heisenberg** developed the uncertainty principle concerning an electron's location and velocity. Electrons are added one at a time to the lowest energy levels first (**Aufbau Principle**). An orbital holds a maximum of two electrons (**Pauli Exclusion Principle**). Electrons occupy equal-energy orbitals so that a maximum number of unpaired electrons results (**Hund's Rule**). Energy levels are designated 1–7. Orbitals are designated s, p, d, and f according to their shapes(sphere, dumbbell, 4-leaf clover.) The s, p, d, f orbitals relate to regions of the Periodic Table. **Valence electrons** occupy the highest principle energy level of an atom. All the elements in a group have the same number of valence electrons. An element's electron configuration determines the number of valence electrons. Example: Bromine's valence electron configuration is $4s^2 4p^5$ with 7 valence electrons. The outermost electrons in an atom are called **valence electrons**. The period (row) number on the periodic table corresponds to the outermost energy level occupied by the valence electrons in an element. Elements in the same group (column) on the periodic table have the same number of valence electrons **Lewis dot diagrams** show the valence electrons of an atom. The electrons (dots) are arranged around the element's symbol. **Metallic bonds** consist of the attraction of free-floating valence electrons for the positively charged metal ions.	**Ch 13: Read pp 361-364** on development of modern quantum mechanical model. **Ch 13: Read pp 364-366** on orbital shapes. **Ch 13: Read pp 367-369** on electron configurations and Aufbua principle. **Ch 13: Read pp 372-381** on the relationship between atomic orbitals and atomic emission spectra. **Ch 15**: Read pp 413-424 on ionic compounds. Read pp 427 on metallic bonds.
Nomenclature, Formulas, and Reactions **3.3** **SOL 3a, 3d, 3e**	Bonds form between atoms to achieve stability. Ionic compounds are formed by the attraction between positive and negative ions. Ions are formed by **electron transfer** from a metal to a non-metal. After electron transfer, both ions meet the **octet rule**. The octet rule is the tendency of an atom to take on the configuration of a noble gas.	**Ch 6**: Review pp 146-156 for naming and writing formulas for ionic compounds.
Molar Relationships **4.4** **SOL 4d**	Dissolving is a physical change that involves heat. When an Ionic compound dissolves in water it breaks into the ions that make it up. This process is called **dissociation** and can be expressed by an equation. Example: $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ Ionic compounds that dissociate completely in water are strong electrolytes.	**Ch 17:** read pp 483-485
Phases of Matter and Kinetic Molecular Theory **5.3** **SOL 5d,5e,5f**	**Specific Heat Capacity**(C) is a physical property of a substance. $\boxed{Q = mC\Delta T}$ is use to calculate heat, mass, specific heat or temperature change respectively. Specific heat can be used to identify a substance. Atoms and molecules are in constant motion. Forces of attraction between molecules determine the physical changes of state. The intermolecular forces must be overcome in order for a substance to melt or boil. Phase changes that require heat (like melting or boiling) are **endothermic**. ΔH is positive for an **endothermic** change. This means heat goes in. **Molar heats of fusion and vaporization** can be used to calculate energy changes. Phase changes that give off heat (like freezing and condensing) are **exothermic**. ΔH is negative. This means heat is released. **Molar heats of solidification and condensation** can be used to calculate energy changes. Heating and cooling curves, known as **temperature line graphs,** show the energy changes that occur as a substance goes from a solid to a gas as temperature is changed.	**Ch 11:** read pp 293-302 on heat capacity and specific heat capacity. Read pp 307-311 on molar heats of phase changes. Review the temperature line graph in figure 11.15 on page 310.

Objectives for Unit 4

Chemistry, Addison-Wesley, 2002

Topic Outline

I) Thermochemistry Part 1(Chapter 11)
  A) Types of Energy
  B) Exothermic and Endothermic processes
  C) Heat Capacity and Specific Heat (p 295-299)
    1) Calculations using specific heat capacities (p.299: 1-3, 8-10)
    2) Calorimetry (p. 300-306)
  D) Changes of State and Heat Changes (p. 307-313)
    1) Phase Changes and Interpreting Heating Curves
    2) Molar Heats of Fusion and Solidification (p.309: 20, 21)
    3) Molar Heats of Vaporization and Condensation (.311: 22, 23)

II) Electrons in Atoms (Chapter 13)
  A) Review Rutherford's Model
  B) Bohr's Model
  C) Quantum mechanical model (Schrodinger & Heisenburg)
  D) Atomic electron orbitals (s,p,d & f) and electron configurations
  E) Identifying valence electrons

III) Ionic Bonding (Chapter 15)
  A) Valence Electrons (read p413-414, Problems p 418#1,3)
  B) Octet Rule (read p414-418, Problems p418#2,4, 5,6)
  C) Ionic Bonding (read p419-421, p 421#7)
  D) Properties of Ionic Compounds (p422-425, Problems p 425#9-13)
  E) Properties of Metallic Bonds and Metals (p427-428, Problems p429#15,17;)

(SOL) Learning Objectives

1. (4e) Identify a process as endothermic or exothermic based on whether it absorbs or releases heat.
2. (5f)Memorize and use $q=mC\ \Delta T$ to solve specific heat capacity and calorimetry problems.
3. (5e) Calculate energy changes during phase changes using molar heat of fusion, molar heat of vaporization,
4. (5d) Identify freezing point, ΔH_{fusion}, $\Delta H_{vaporization}$, and boiling point on a heating curve of water.
5. (2f) Determine the # of valence electrons and electron configurations for anions and cations
6. (3d) Explain why ionic bonds form in terms of electron transfer and the octet rule
7. (3d) Explain why Hydrogen, Lithium, and Beryllium break the octet rule in ionic compounds
8. (3d) Predict which compounds will be ionic based on their position on the periodic table.
9. (2h) Define an electrolyte
10. (2h) Predict which compounds will be electrolytes
11. (2h) Illustrate what happens when an ionic compound dissolves in water.
12. (2h) Explain why metals conduct electricity
13. Identify the contributions of Bohr, de Broglie, Planck, Heisenberg and Shrodinger to the development of the modern atomic model.
14. Use the Pauli Exclusion Principle, the Aufbau Principle, and Hund's Rule to determine electron configurations.
15. Identify the shapes of the s, p and d orbitals and the number of electrons in each.
16. Provide the spdf orbital electron configuration of elements using the periodic table.

Chapter 11 Noteguide Part 1: Energy and Changes of State

I.ENERGY CHANGES

A. Definitions

Energy: capacity to do work (movement), or supply heat

Potential Energy: stored energy (food, gasoline, wood... any fuel)

Kinetic Energy: energy due to motion $KE = \frac{1}{2}mv^2$ (like "speed")

Heat: A form of Energy... transferred from "hot" to "cold" the measure of heat = temperature.

Thermochemistry: Study of heat changes from chemical reactions or phase changes $\boxed{\Delta T}$

B. Exothermic and Endothermic Processes

heat released heat absorbed

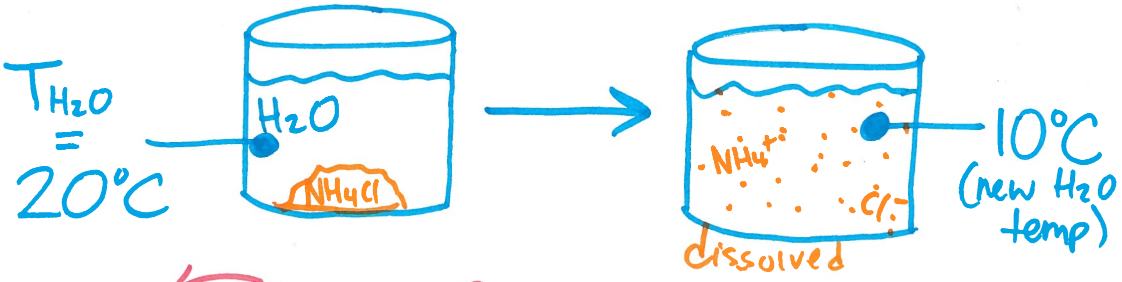
Processes that absorb (endo) heat or release (exo) heat.

Definitions

Consider $NH_4Cl(s) \xrightarrow{water} \overbrace{NH_4^+(aq) + Cl^-(aq)}^{dissolving}$

ammonium chloride

- NH_4Cl dissolved, and ∴ absorbed energy from H_2O (endothermic)



Enthalpy: Synonym for "heat" heat in a system. (at constant pressure)

System: The part of the universe being studied (NH_4Cl)

Surroundings: Everything else...!

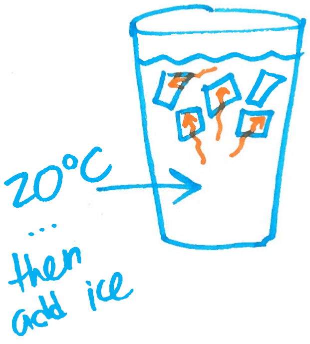
ΔH: "Change in heat" or Δ enthalpy (in the system)

Endothermic Process: heat flows INTO system $\boxed{+\Delta H}$

Exothermic Process: heat flows OUT OF system $\boxed{-\Delta H}$

$E=mc^2$
↑ ↑

Another example of Endothermic and Exothermic processes
ice melting in glass of water:



System: $\boxed{\text{ice cubes}}$ (@ -5°C) $+\Delta H$

Surroundings: water (@ 20°C) $-\Delta H$

(heat into ice)

ice melts (warms)

water cool

$$\boxed{H_2O(s) + heat \rightarrow H_2O(l)}$$

$$H_2O(s) \xrightarrow{\Delta} H_2O(l)$$

Law of Conservation of Energy: Energy is neither created nor destroyed... only TRANSFERRED!

II. HEAT CAPACITY AND SPECIFIC HEAT CAPACITY

A. Definitions

Joule (J): SI unit for ENERGY   one kJ = 1000 J   1000

1 calorie: Amount of energy needed to raise the temp. of 1 g of H_2O by 1°C

1000 calorie = 1 Calorie (food labels) and 1 calorie = 4.184 J

g/L

Specific Heat Capacity: Amt of heat needed to raise the temp of 1 g of a substance by 1°C

Symbol: C   Units: $\boxed{\frac{J}{g \cdot {}^\circ C}}$ or $\frac{J}{g \cdot K}$

water holds heat

Equation:

q = heat

C= specific heat capacity   m= mass

ΔT= change in... temperature

$T_i = 21^\circ C$ $T_f = 12^\circ C$

$\Delta T = -9^\circ C$

$q = mC\underline{\Delta T}$

$\Delta T = T_2 - T_1$

$\Delta T = T_f - T_i$

B. Solving Specific Heat Capacity Problems

$$q = m\,C\,\Delta T$$

The equation has four variables: "q" is heat in Joules; "m" is mass in grams; "C" is specific heat capacity in J/(g•°C); "ΔT" is change in temperature in °C (the change in temperature is the final temperature minus the initial temperature, or $\Delta T = T_f - T_i$). This equation is only valid if the substance does not change phases. Identify the variables, then solve for the missing variable.

Sample Problems

$\Delta T = 35.0 - 22.0 = +13.0$

$C = 0.46$ J/g·C

1. A 500 $\overset{m}{\text{g}}$ sample of iron changes from 22.0°C to 35.0°C. The specific heat of iron is known to be 0.46 J/(g•°C). How much heat was added?

$q = mC\Delta T$

$= (500)(0.46)(+13)$

$q = 2{,}990$ J

2. A 500. g sample of water changes from 22.0° to 35.0°C. The specific heat of water is known to be 4.18 J/(g•C) How much heat was added?

$q = mC\Delta T$

$q = (500)(4.18)(+13)$

$= 27{,}170$ J

$\Delta T = +15.3$°C

3. When $\overset{q}{82}$ J of heat is added to a sample of aluminum, its temperature increased by 15.3°C. Given that the specific heat capacity of aluminum is $\overset{c}{0.90}$ J/(g•°C), what is the mass of the sample?

$q = mC\Delta T$

$C\Delta T \qquad C\Delta T$

$\Delta T = T_2 - T_1$

$m_{(grams)} = \frac{q}{C\cdot\Delta T} = \frac{82}{(0.90)(15.3)} = 6.0$ g Al

4. It takes 78.2 J to raise the temperature of 45.6 g lead by 13.3°C. Calculate the specific heat capacity of lead.

5. Challenger: A $\overset{m}{142\text{ g}}$ sample of silver at a temperature of $\overset{T_1}{19.6}$°C absorbs $\overset{+}{61.30}$ J of heat. What is the final temperature of the sample? [C_{Ag} = 0.24 J/(g•°C)] (*Ans = 21.4°C*)

$q = mC\Delta T$

$\Delta T = T_2 - T_1$

$q = mC(T_2 - T_1)$

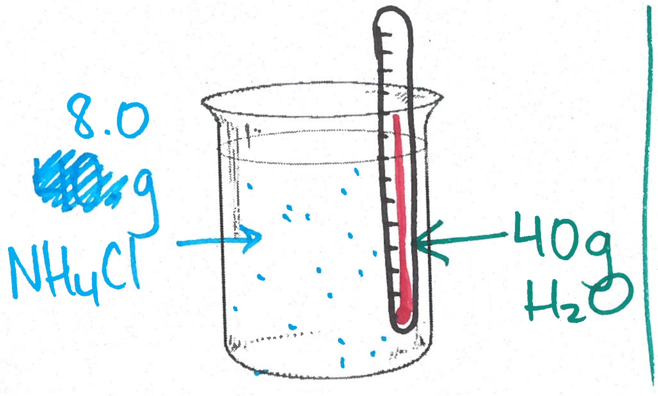
$\frac{61.30}{(142\cdot 0.24)} + 19.6 = T_2 = 21.4$°C

$\frac{q}{mc} = T_2 - T_1 \rightarrow \frac{q}{mc} + T_1 = T_2$

C. Calorimetry

Calorimetry: The msmt of heat change for a chemical or physical process (food; fuel...)

Calorimeter: Device for measuring heat change.



$m_{H_2O} = 40.\,g$

$C_{H_2O} = 4.184 \frac{J}{(g\cdot°C)}$

$\Delta T = -10°C$

$q = mc\Delta T$

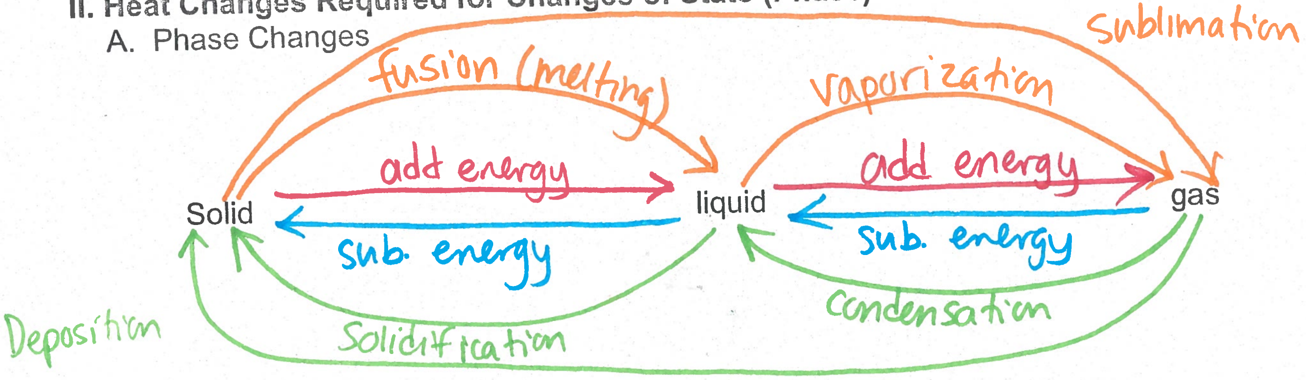
$q = -1673.6J$

$-q_{H_2O} = q_{NH_4Cl}$

$-1673.6J = \boxed{+1673.6\,J}$

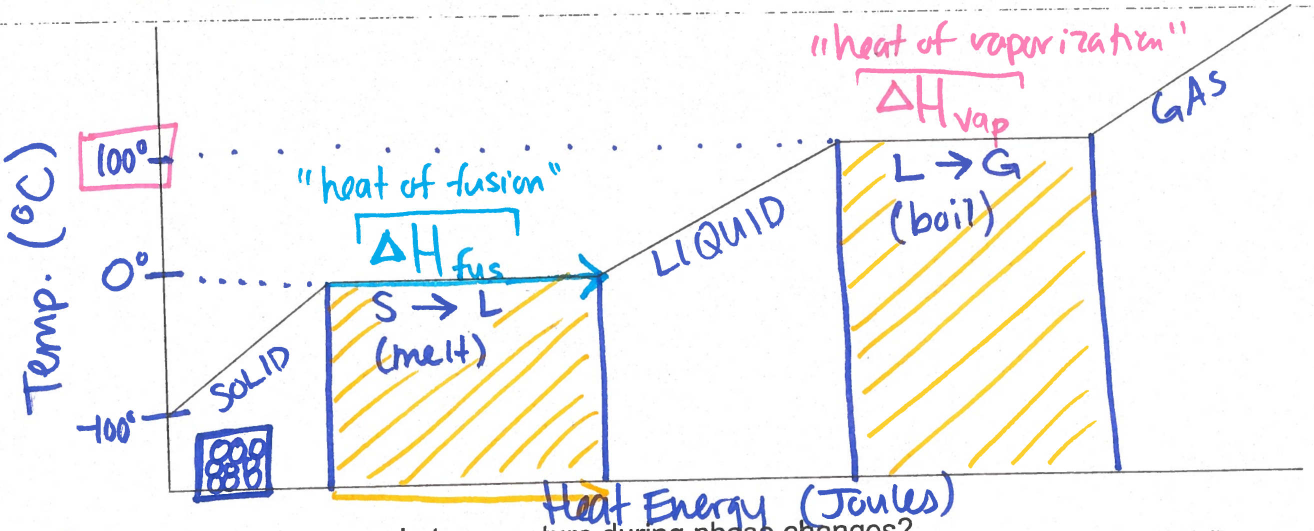
II. Heat Changes Required for Changes of State (Phase)

A. Phase Changes



Heat is absorbed or released when matter changes state.

Interpreting Heat Curve for water Relate T to ΔH, with phases



Does water increase in temperature during phase changes?

No... all Ⓔ goes into overcoming intermolecular attractions melting

11/7

B. Molar Heats of Phase Changes

Molar heat of fusion: ΔH_{fusion} = heat absorbed ~~(or released)~~ by 1 mole of a melting solid

J/g or J/mol or cal/g or cal/mol

Molar Heat of Solidification: $\Delta H_{solid.}$ heat released when 1 mol of something (Liquid) solidifies → (same possible units)

$$\boxed{\Delta H_{fusion} = -\Delta H_{solidification}}$$

Molar heat of vaporization ΔH_{vap} : Amt of heat absorbed by L→G transition

Molar Heat of condensation: $\Delta H_{cond.}$ opp. (exo)

$$\boxed{\Delta H_{vap.} = -\Delta H_{cond.}}$$

Molar Heats apply to phase changes The units may include:

J/mol J/gram calories/gram

KJ, or J, or cal / mol or grams.

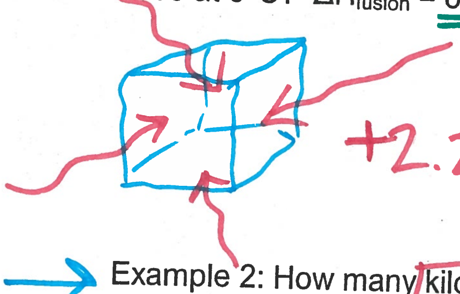
ENERGY/AMOUNT

Solve the problems as unit analysis problems

C. Using Molar Heats in Calculations.

$H_2O(s) + HEAT \rightarrow H_2O(l)$

Example 1: How many grams of ice would be melted by adding 2.25 kJ of heat to an ice cube at 0°C? ΔH_{fusion} = 6.0 kJ/mol



+2.25 KJ

$$\frac{2.25\ \cancel{KJ}}{1} \times \frac{1\ \cancel{mol\ ice}}{6.0\ \cancel{KJ}} \times \frac{18.0\ g\ H_2O}{1\ \cancel{mol}\ ice}$$

$$6.75\ g\ H_2O(s)$$

Example 2: How many kilojoules of heat would be released when 36.04 grams of steam condenses to water at 100˚C?

$$\boxed{H_2O(g) \rightarrow H_2O(l) \text{ and } \Delta H_{cond.} = -40.7\ kJ/mol}$$

exothermic

steam ↓ -ΔH liquid H_2O

$$\frac{36.04\ \cancel{g\ H_2O(g)}}{1} \times \frac{1\ \cancel{mol\ H_2O}}{18.0\ \cancel{g\ H_2O}} \times \frac{-40.7\ KJ}{1\ \cancel{mol}}$$

$$\Delta H = -81.49\ KJ$$

"released" = exo.

Energy/amount

Molar Heat Calculations Practice

molar heat is given in J/mol or J/g or cal/g, so use it as a conversion factor

1. How much heat is required to melt 500.9 grams of ice at 0°C? The heat of fusion of water is 80.0 cal/g.

$$\frac{500.9\ \cancel{g\ H_2O(s)}}{1} \times \frac{80.0\ cal}{1\ \cancel{g}} = \boxed{40,072\ cal}$$

2. How much heat is required to vaporize 13.1 grams of methane (CH_4) at it's boiling point, which has a heat of vaporization of 8.2 kJ/mol?

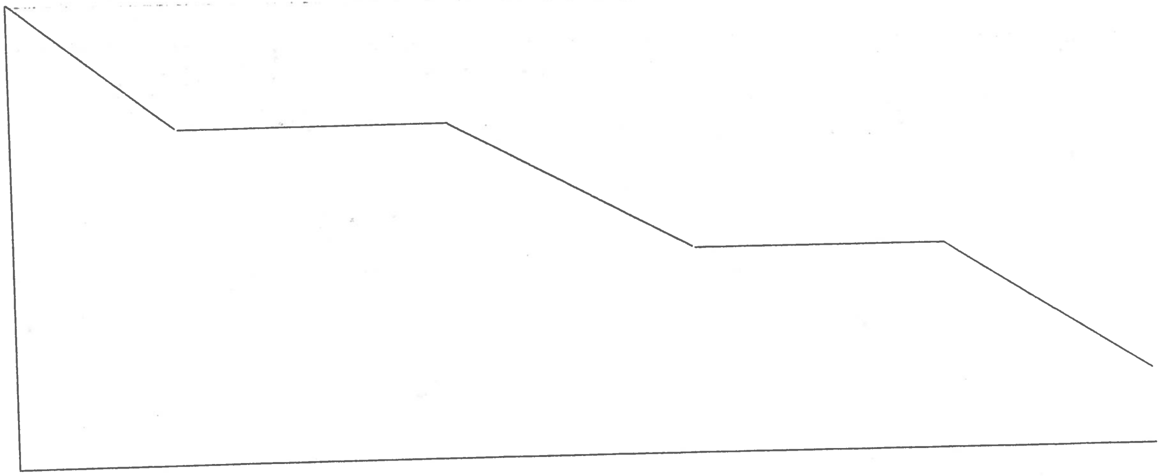
$$\frac{13.1\ \cancel{g\ CH_4}}{1} \times \frac{1\ \cancel{mol\ CH_4}}{16.0\ \cancel{g\ CH_4}} \times \boxed{\frac{8.2\ kJ}{1\ \cancel{mol}}} =$$

3. How many grams of neon must crystallize (solidify) at it's freezing point to release 560 J of heat, given that the neon's ΔH_{fusion} = 330 J/mol?

$$\frac{560\ J}{1} \times \underline{\qquad\qquad} \times \underline{\qquad\qquad}$$

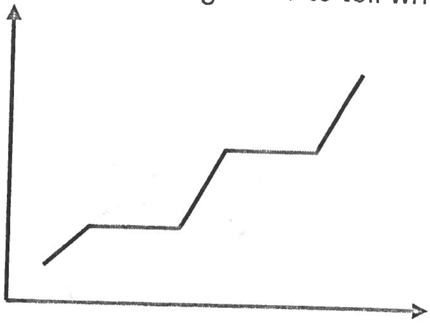
$\frac{11/7}{11/9}$

Interpreting a Cooling Curve for Water



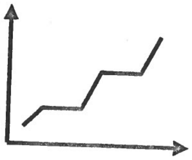
Mixed Molar Heats and Specific Heat Capacity Problems

Use the heating curve to tell which is which

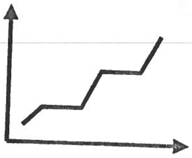


Specific heat capacity $H_2O(s)$ = 2.1 J/(g°C)
Specific heat capacity $H_2O(l)$ = 4.2 J/(g°C)
Heat of fusion H_2O = 6.0 kJ/mol
Heat of vaporization H_2O = 41 kJ/mol

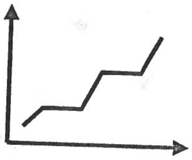
1. How much energy is needed to raise the temperature of 150 grams of ice from -20.0°C to -5.0°C? *(Ans = 4725 J)*



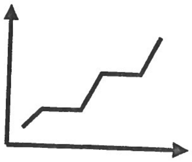
2. How much energy is needed to vaporize 52 grams of water at 100°C? *(Ans = 118 ≈120 kJ).*



3. How many grams of ice at 0°C would be melted by adding 820 kJ of heat. *(Ans = 2500 g ice)*

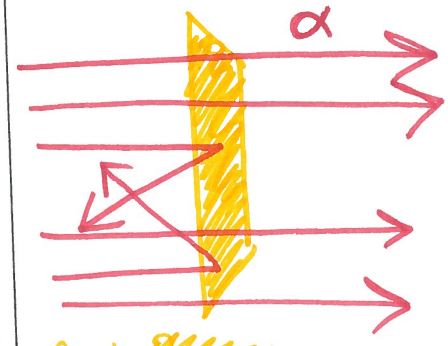
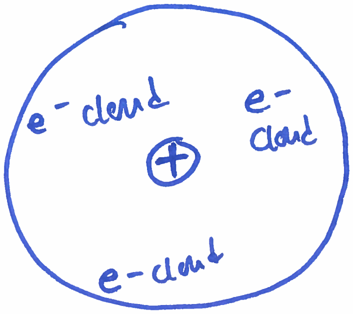


4. How much will the temperature of 850 grams of water increase if 16,000 Joules of heat is added? *(Ans = 4.5°C)*



Chapter 13: Electrons in Atoms

Review of Rutherford's Atomic Model(1911)

What	*How*	*Model*
Atoms are mostly empty space... with ⊕ly-charged, tiny nucleus		

Gold ~~[illegible]~~ Foil exp.

A. Important Terms

1. atomic number: number of protons-whole number shown on the periodic table
2. mass number : number of protons plus neutrons Carbon-13 → 6 p^+ → 7 n^0
3. isotopes: elements with the same number of protons, but a different number of neutrons
4. atomic mass: weighted average of isotope masses. Listed on the periodic table.

6
C
12.01

B. Symbols for Isotopes neutral

1. $^{13}_{6}C$ 6 protons, 6 electrons, 7 neutrons
2. ^{64}[redacted]Cu 29 protons, 29 electrons, 35 neutrons Copper-64
3. Pb-202 82 protons, 82 electrons, 120 neutrons

p+n

C. Practice

OXO

ISOTOPE	ATOMIC #	# PROTONS	# NEUTRONS	MASS #
^{54}Fe				
	36		40	
		13	14	

OXO

How many electrons, neutrons and protons in Zinc-67?

How many neutrons are in F-19?

An Aside About Light and Energy

← 300,000,000 m/s

Light is fast. It travels at 3.0×10^8 m/s. (*distance* over *time* is *speed*, which is the magnitude of velocity). "c" is the constant that represents light's speed in a vaccuum.

or "f" Light **frequency** (ν, called "nu") × Light **wavelength** (λ, called "lambda") = c

Max Planck (1900) determines Said "Energy of light is directly proportional to its frequency" $E \propto \nu$

Equations:

Energy of light, using frequency:

$$E = hf = h\nu$$

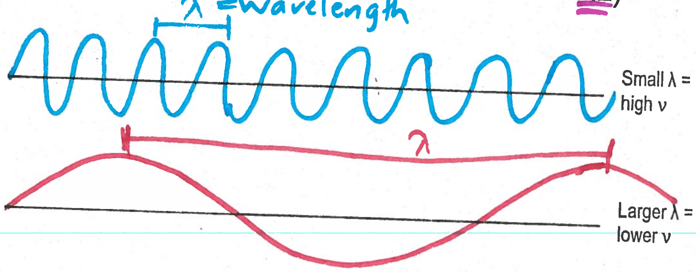
Energy of light, using wavelength:

$$E = h\frac{c}{\lambda}$$

• = constants

h = 6.626 × 10^{-34} J·s

ν ("nu" not "v") = frequency (unit: 1/s ≡ s^{-1} ≡ Hz)



Red	Orange	Yellow	Green	Blue	Indigo	Violet

Wavelength = 7.0 x 10^{-7} m
Low Energy

wavelength
4 x 10^{-7} m
High Energy

Louis de Broglie (1924) determines All matter has wave-like properties. SMALL particles (i.e., e^-, p^+, n^0) have BIG waves. BIG have SMALL waves

The de Broglie Equation & Interpretation:

$$\lambda = \frac{h}{m \cdot v}$$

v = velocity

h = planck's constant: 6.63×10^{-34} J·s

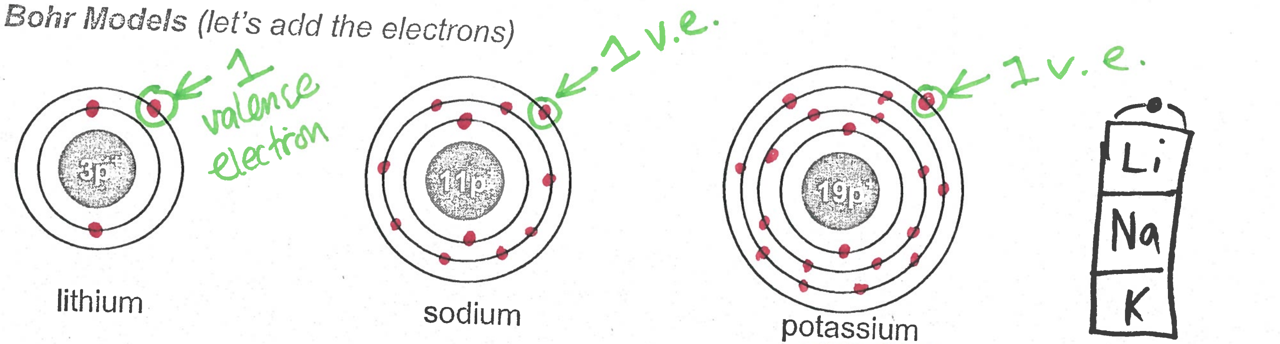
What did the spectrum tube demonstration show?

All elements have unique electron configurations, ∴ unique spectra.

Bohr's Model of the Atom (Powerpoint)

What	*How*	*Model*

A Few Bohr Models (let's add the electrons)



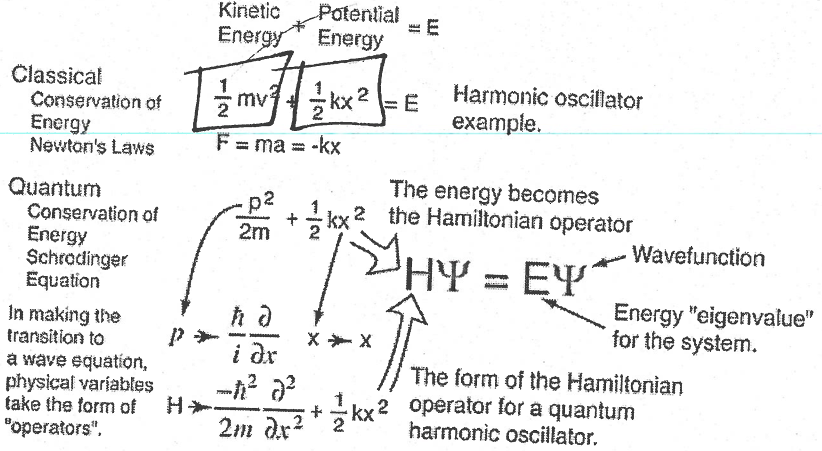
The number of electrons in the outer principal energy level (or valence shell) is the same within a group. same family

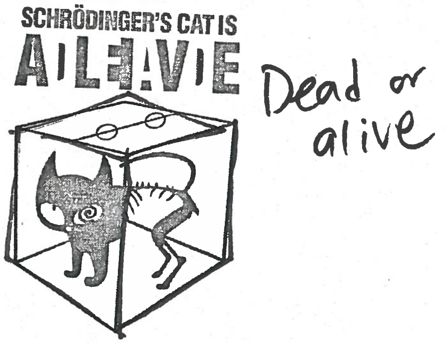
The number of principle energy levels is the same as the period #.

So why isn't this model good enough?

The Quantum Mechanical Model

A. Erwin Schrödinger used complex mathematics to calculate where electrons **probably** are around the atom. His mathematical models were revolutionary to physics.





B. Werner Heisenberg adds the **Heisenenberg Uncertainty Principle:**

Can't know position & velocity at same time.

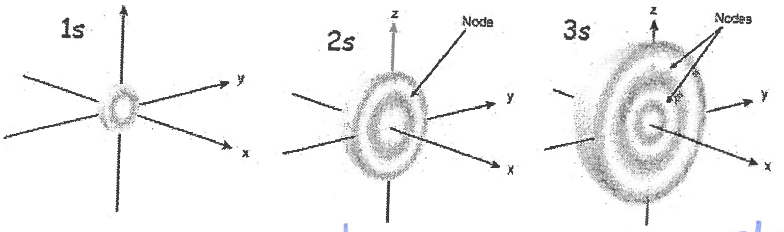
Know 1 or other.

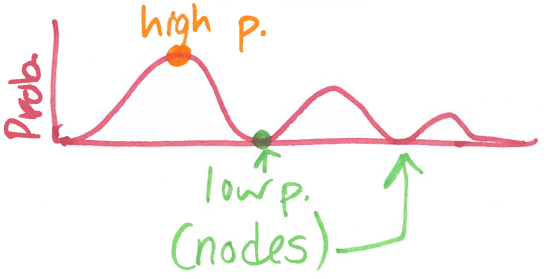
$$\Delta x . \Delta p \geq \frac{h}{4\pi}$$

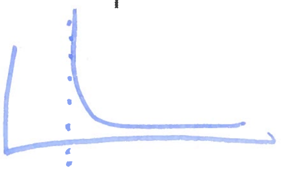
$$\Delta p = m \, \Delta v$$

$$\boxed{\Delta x . m\Delta v \geq \frac{h}{4\pi}}$$

C. The areas where an electron can *probably* be found are called orbital Clouds. (The areas where electrons will be unlikely are called nodes)





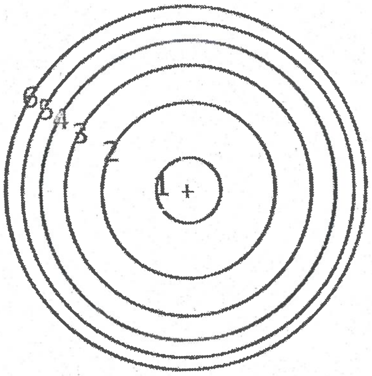


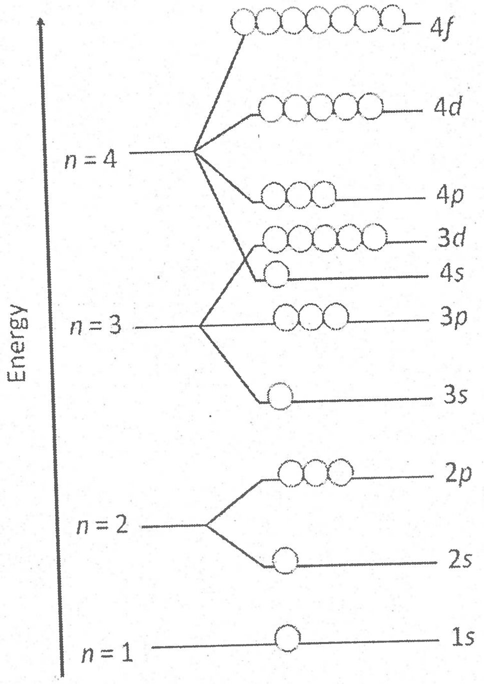
D. Each orbital has a specific shape and can hold up to 2 e^- (spinning in opposite directions).

a. How is the spin of an electron noted in models? using arrow

↿⇂ or ↑ ↓

E. Organization of Electrons



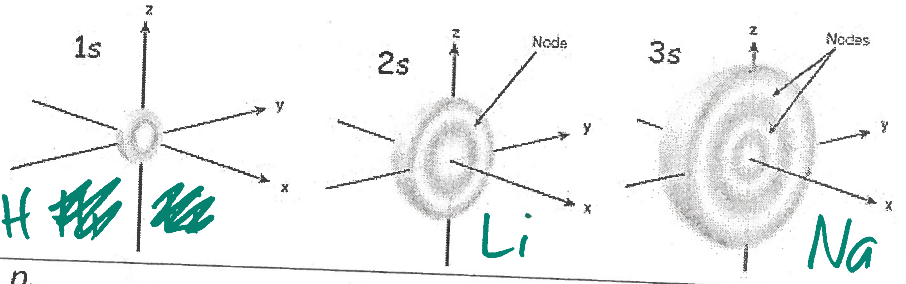
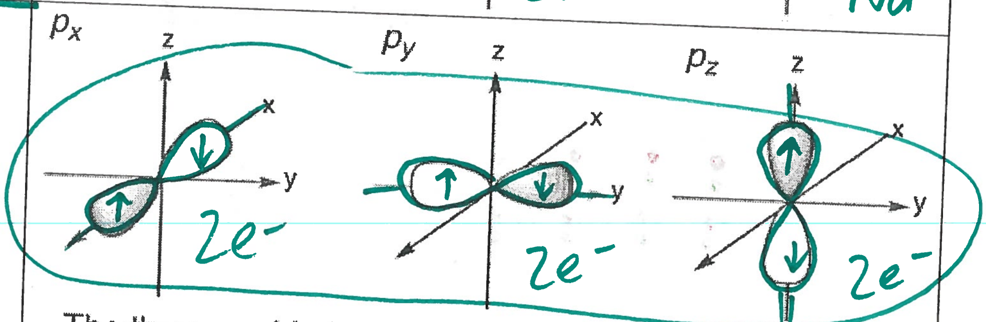
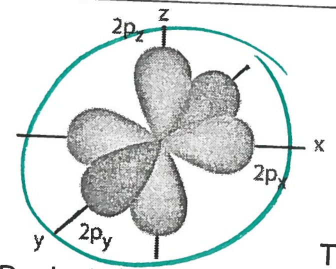
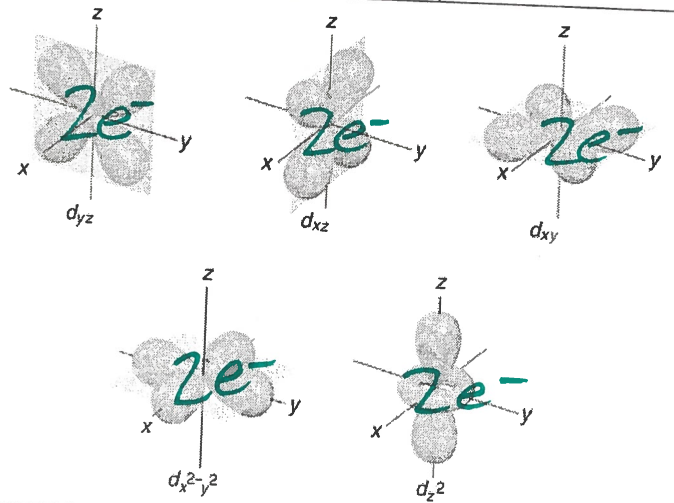


Description of Sublevels

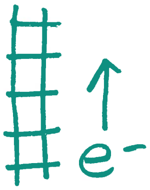
1. "s" sublevels have 1 orbital – it is spherical shaped
2. "p" sublevels have 3 orbitals – they are dumbell / hourglass shaped
3. "d" sublevels have 5 orbitals – 4 are cloverleaf shaped, one "pacifier"
4. "f" sublevels have 7 orbitals – they are Crazy shaped

11/16

11/17

What orbital type is this? S How many electrons can go in this orbital? 2 e⁻ each.	
How many electrons can fill each p- orbital? 2 ~~e~~ e⁻	 The three p orbitals are aligned along perpendicular axes
How many electrons can fill this entire p-energy sublevel? 6 e⁻	 6e⁻ The p-energy sublevel is made of all three 3D orientations (p_x, p_y, and p_z together)
How many electrons can fill each d-orbital? 2e⁻ How many electrons can fill this d-energy sublevel? 10e⁻	

F. Filling in the Orbitals in Quantum Mechanics



1. **Aufbau Principle**: "to build up"

Electrons fill the lowest energy levels first. (start @ "1s")

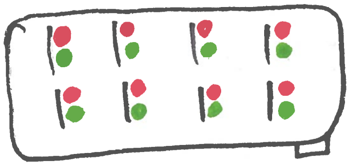
2. **Pauli Exclusion Principle** (PEP): ↑↓ or ⇅

Only $2e^-$ electrons can be in each orbital (two per box or line!) and they must have opposite magnetic spins. (two different arrow directions!)

$\boxed{\uparrow\downarrow}$ or ↑↓

3. **Hund's Rule**:

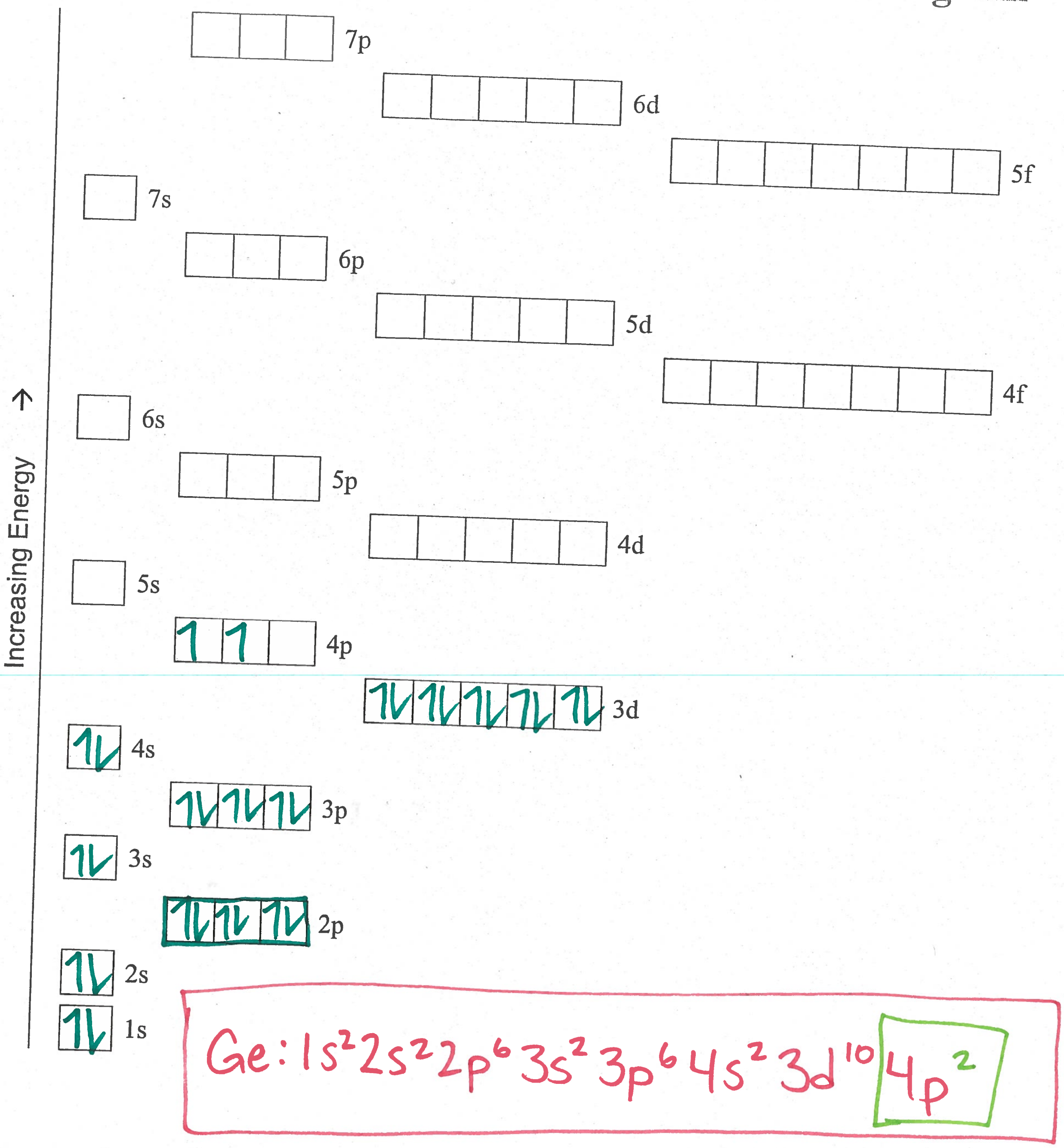
When electrons occupy orbitals of equal energy, they fill in singly with aligned spins *before* they double up (space out if you can!) The bus seat analogy…



Bus

Germanium 32 e^-

Electron Configurations & the Aufbau Diagram

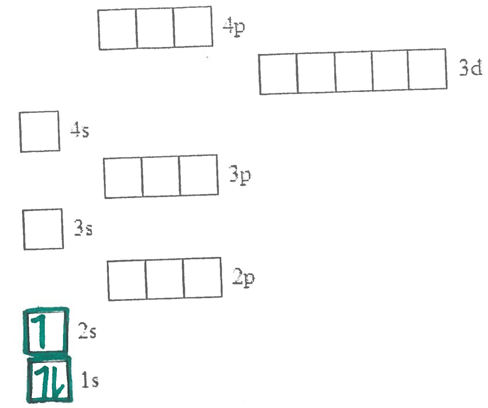


$Ge: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} \boxed{4p^2}$

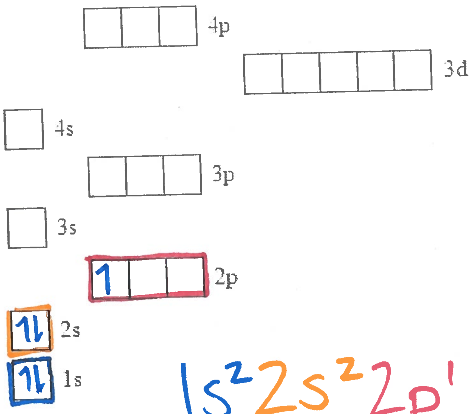
Rules to fill it in:

✓1. Electrons enter lowest energy first. (start with "1s") *[Aufbau Principle]*
✓2. An orbital can have at most 2 electrons with opposite spins. *[Pauli Exclusion Principle]*
✓3. When electrons are filling orbitals of equal energy, one electron enters each before they start to spin pair (double up). *[Hund's Rule]*

Li #electrons = $3e^-$

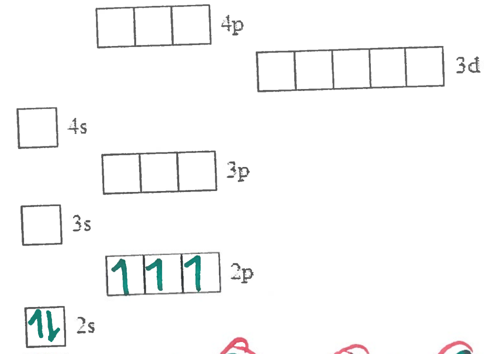


B # electrons = $5e^-$



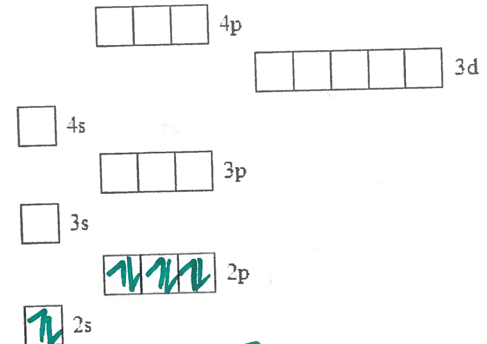
$1s^2 2s^2 2p^1$

N # electrons = $7e^-$



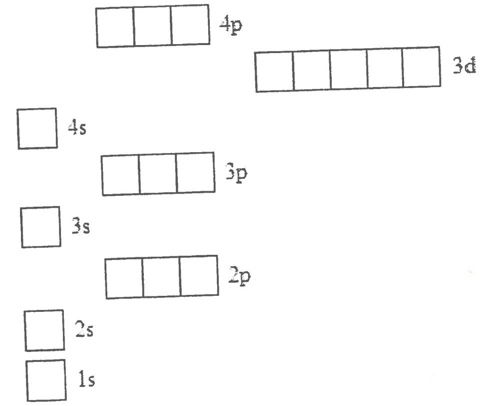
$1s^{\textcircled{2}} 2s^{\textcircled{2}} 2p^{\textcircled{3}}$

Ne # electrons = $10e^-$

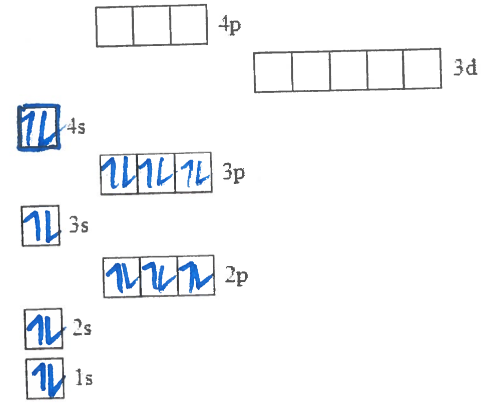


$1s^2 2s^2 2p^6$

P # electrons =



Ca #electrons = $20e^-$



oYo **As** # electrons = 33e⁻

↑ ↑ ↑ 4p

↑↓ ↑↓ ↑↓ ↑↓ ↑↓ 3d

↑↓ 4s

↑↓ ↑↓ ↑↓ 3p

↑↓ 3s

↑↓ ↑↓ ↑↓ 2p

↑↓ 2s

↑↓ 1s

oYo **Cr** # electrons =

☐☐☐ 4p

☐☐☐☐☐ 3d

☐ 4s

☐☐☐ 3p

☐ 3s

☐☐☐ 2p

☐ 2s

☐ 1s

Complete the alternate form of the Aufbau diagrams below:

14e⁻ Si	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑							
	1s	2s	2p	2p	2p	3s	3p	3p	3p	4s	3d	3d	3d	3d	3d
Mg															
	1s	2s	2p	2p	2p	3s	3p	3p	3p	4s	3d	3d	3d	3d	3d
Fe															
	1s	2s	2p	2p	2p	3s	3p	3p	3p	4s	3d	3d	3d	3d	3d

Indicate how many unpaired electrons each element has

Si: 2 e⁻

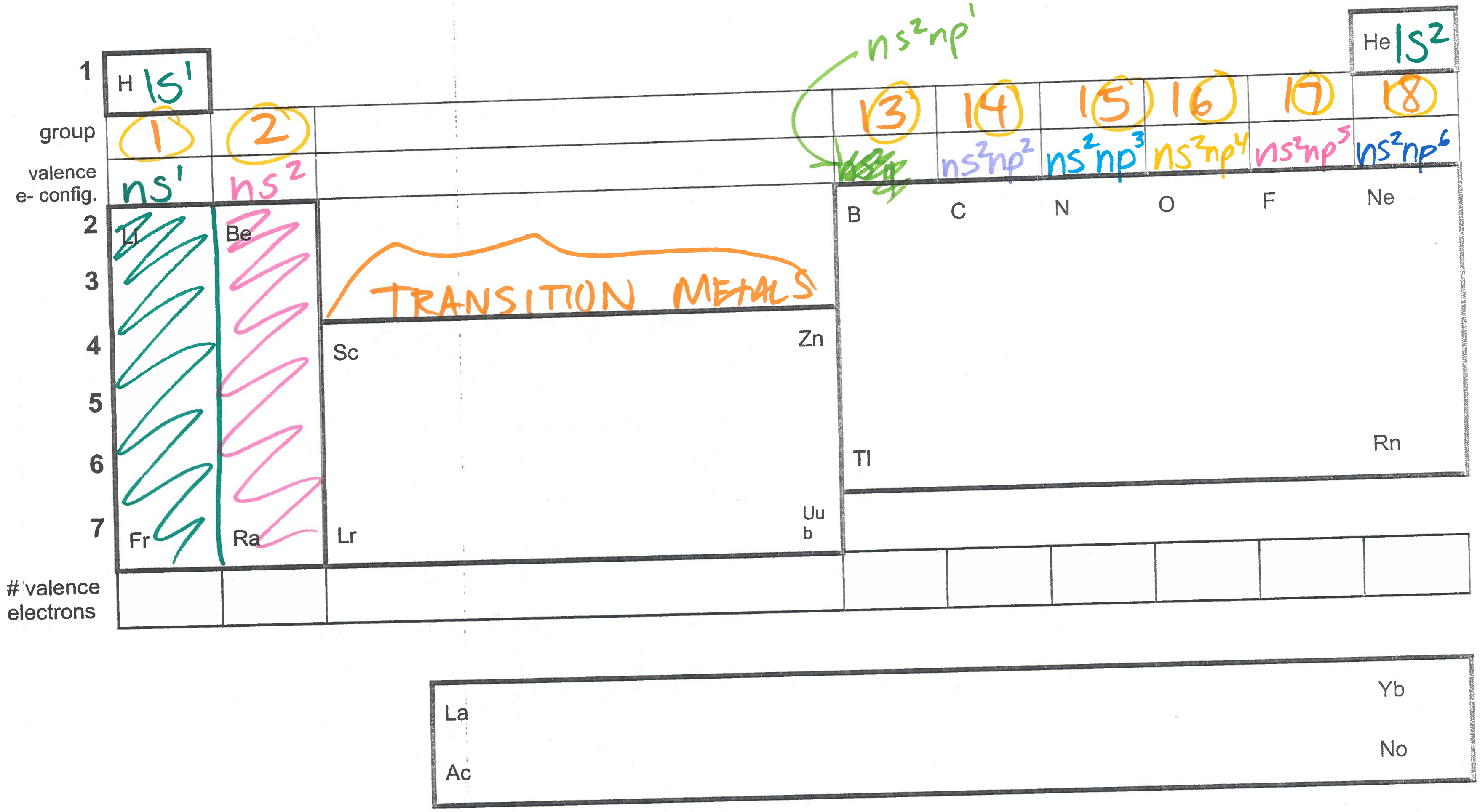
Mg:______

Fe:______

valence electrons: $ns^x np^y$

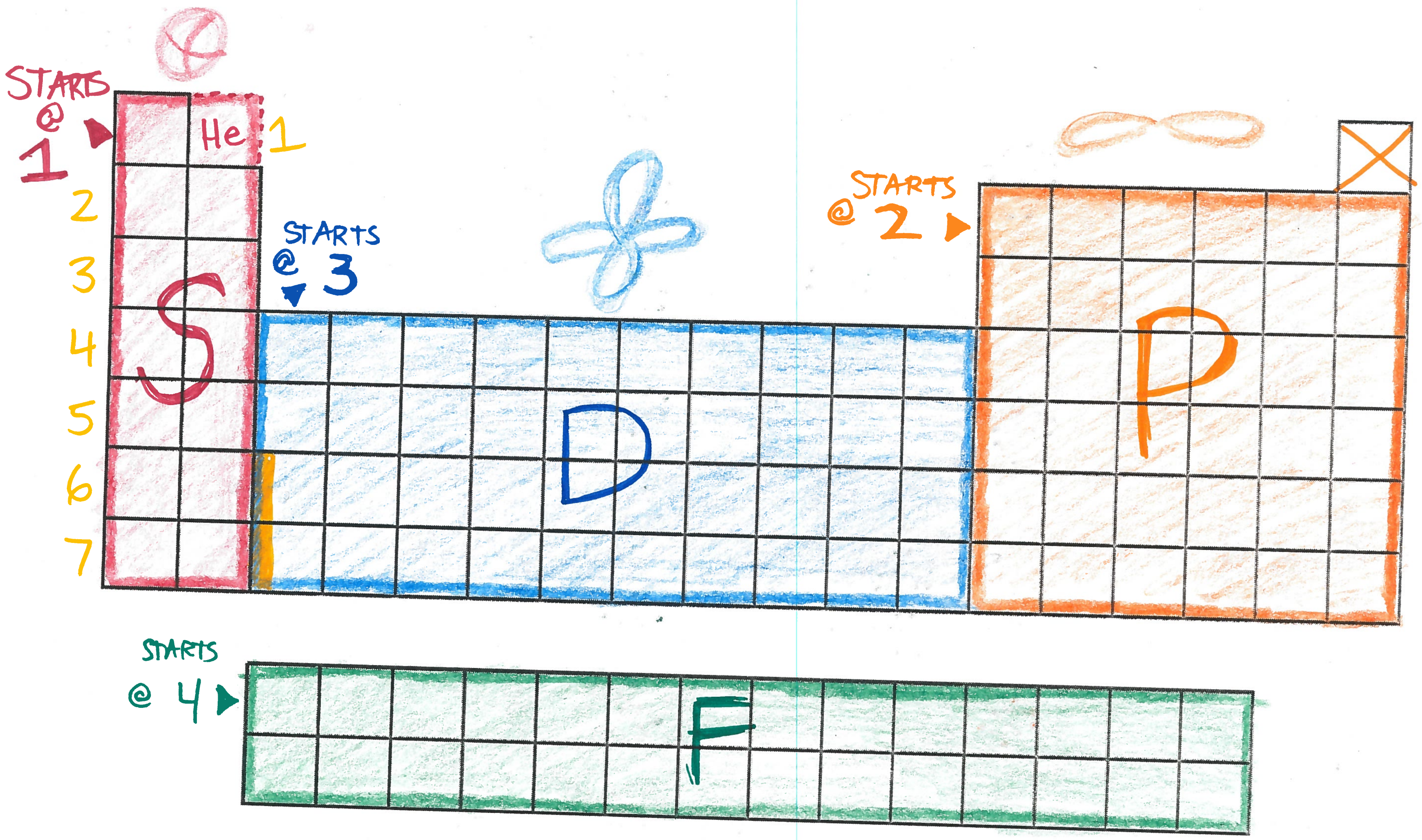
$x + y = \#$ of valence electrons

Chapter 13: Periodic Table and Electron Configuration



Representative elements: "S" & "p" blocks

TALL COLUMNS


STARTS @ 1
He
1
2
3
4
5
6
7
S
STARTS @ 3
D
STARTS @ 2
P
STARTS @ 4
F

Chapter 15: Ionic Bonds

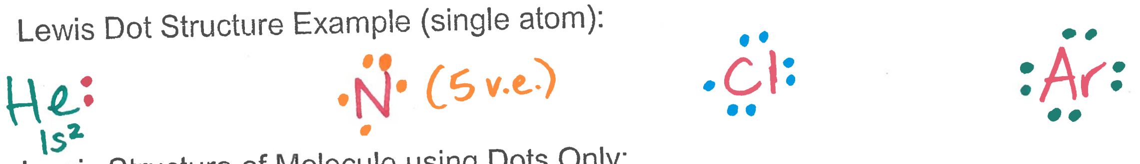
I. **Valence** Electrons

A. definition: e- in outermost shell (highest energy) The ones that bond!

B. Lewis Dot Structures: show valence electrons as dots; the symbol represents the core electrons (which is everything *but* the valence electrons).

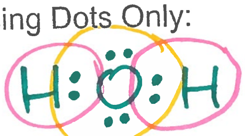
Lewis Structures show bonded atoms as lines.

Lewis Dot Structure Example (single atom):

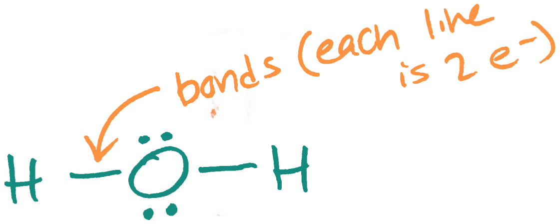


Lewis Structure of Molecule using Dots Only:

H_2O



Lewis Structure of Molecule using Lines for Bonds:



Group	1	2	13	14	15	16	17	18
Example	Na	Ca	B	Si	P	O	Br	Kr

II. **Octet** Rule 8

A. definition: atoms are generally most stable with 8 v.e. (tries to look like nearest noble gas)

B. A full valence shell is very stable (which is happy! ☺). Therefore, elements gain or lose electrons to reach a full octet.

$ns^x np^y$

- configuration example: Na = $1s^2 2s^2 2p^6 \boxed{3s^1}$

Na· lost

Na⁺ = $1s^2 2s^2 2p^6$

- configuration example: S = $1s^2 2s^2 2p^6 3s^2 3p^4$

S²⁻ = $1s^2 2s^2 2p^6 3s^2 3p^6$

C. *Exceptions* to Octet Rule in Ionic Compounds

Helium is happy with 2 valence electrons, so we call this exception the duet rule. Atoms with atomic numbers close to He (such as H, Li, and Be) will be happy with 2 electrons. They can't fit 8!

$1s^1$ $1s^2 2s^1$

H^+: = $1s^0$ H^- = $\overset{He}{\boxed{1s^2}}$ Li^+ = $\overset{He}{\boxed{1s^2}}$ Be^{2+} = $\overset{He}{\boxed{1s^2}}$

Terminology: **iso-** means same or one or equal (think: isosceles triangle)
-electronic refers to the number of electrons.

∴, what does isoelectronic mean? Same # of v.e.

Concept Check:

- Are He and Ne isoelectronic with each other? No 2 ≠ 8
- Are O^{2-} (8ve) and Ne (8ve) isoelectronic with each other? Yes $8 \cancel{\neq} = 8$
- Are F^- and Cl^- isoelectronic with each other? Yeah...
- Are Cl^- and S^{2-} isoelectronic with each other? Yes!
- Which noble gas will iodine (NM) become isoelectronic to when an iodine atom is ionized? Xenon
- Na+ will lose one electron, to become isoelectronic with Neon.
- Which alkaline metal is most likely to ionize to become isoelectronic with the noble gas Krypton? Sr
- Are Mg^{2+} and N^{3-} isoelectronic? yes 'Ne" = "Ne"

D. A couple of other octet rule **exceptions**:

Boron (B) *actually* prefers to have __________ valence electrons (and it's stable that way!), rather than 8 like many others.

Atoms from sulfur and beyond can sometimes have more than 9. This is called

_______________________.

III. Ionic Bonding

A. Question: *Where do anions get their extra electron(s) from anyway?*

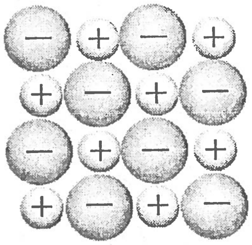
____

____

____

____

Examples: (NaCl, CaF_2, MgO, Li_3P, K_2S)



⊕→ ←⊖

IV. Properties of Ionic Compounds

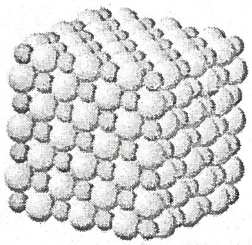
Ionic compounds are held together by Strong Electrostatic attractions

Electrostatic attraction: opp.-charged particles attract to each other.

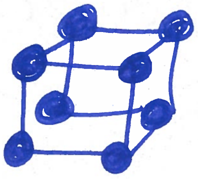
A. crystal structure: 3D orderly repetitive atomic arrangement (lattice)

$C_6H_{12}O_6$

$C_{12}H_{22}O_{11}$



NaCl



B. electrolytes: compounds that conduct electricity when dissolved in H_2O

ionic!

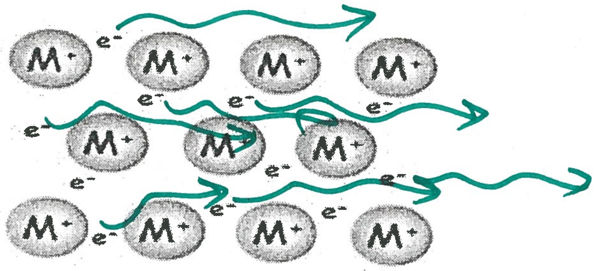
C. high melting points ionic stuff... high MP

V. Metallic Bonds

exist in metals

A. caused by attraction of electrons (valence) for the positively charged nuclei in other atoms.

B. metals are good conductors because of these free floating electrons.



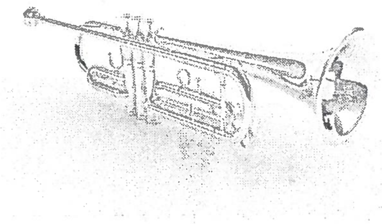
Ever notice how some metals, such as *steel*, *bronze*, and *brass* aren't on the periodic table???

These are called ALLOYS. An **alloy** is a solid mixture of metals.

Two (or more) metals are melted, then mixed together while they're still liquids. After the hot liquid metal mixture cools, you have an alloy.

Brass is made of copper and zinc.

It's great for musical instruments due to how sound waves resonate (propagate) through the metal atoms!



Bronze is made of copper, tin, and other metals.

Steel is made of iron, carbon, and other elements.

Jewelry… what is "white gold" and "rose gold"? Jewelry is often an alloy. White gold is an alloy of gold and another metal, like nickel or platinum. Metallic bonds keep it together, of course.

Investigation Questions:

Why is it not a good idea to have jewelry that is pure gold?

Pure gold is weak, soft, & malleable

Fe

What makes stainless steel special? And why doesn't it stain easily?

Coated with a material that prevents oxidation.

OYO Terms to Know:

Malleable ______________________________

Conductor ______________________________

Ductile ______________________________

Brittle ______________________________

Writing Ionic Formulas from Names Review

SKIP (HW practice)

1. Identify the charge of the cation (see periodic table)
2. Use empty parentheses if you don't know the metal's charge immediately
3. Identify the charge of the anion
4. Identify the charge of the metal by canceling the anion's charges
5. Put the charge of the metal in the empty parenthesis. This is the *oxidation state* of the metal.

Magnesium carbonate

Sodium phosphate

Strontium Nitride

Calcium nitrate

Tin (IV) chloride

Copper (III) Sulfate

Naming Ionic Compounds Review

1. Name the cation
2. Does the cation name need a parentheses
3. Name the anion
4. Figure out the cations charge if needed

Li_2O ______________________________

$CaCl_2$ ______________________________

$Fe(NO_2)_3$ ______________________________

Ba_3P_2 ______________________________

$V(OH)_5$ ______________________________

Cr_2O_3 ______________________________

$Sr_3(PO_4)_2$ ______________________________

$Cu(NO_3)_2$ ______________________________