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:____

Potential Energy:____

Kinetic Energy:____

Heat:____

____

Thermochemistry:____

____

B. Exothermic and Endothermic Processes

Processes that absorb (__________) heat or release (________) heat.
Definitions

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

Enthalpy:____

____

System:____

Surroundings:____

ΔH:____

Endothermic Process:_______________________________________

Exothermic Process:____

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

Law of Conservation of Energy:____

____

II. HEAT CAPACITY AND SPECIFIC HEAT CAPACITY

A. Definitions

Joule (J): ____ one kJ = ________J

calorie: ____

____

_______ 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 ______________ Joules.

Specific Heat Capacity:____

____

Symbol: Units:

Equation:

q = ______________________________

C=______________________________ m=_____________________

ΔT=____

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?

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?

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?

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.

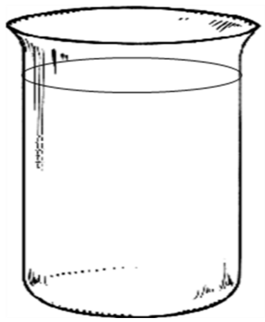
5. A 142 g sample of silver at a temperature of 19.6°C absorbs 61.30 J of heat. What is the final temperature of the sample? [C_{Ag} = 0.24 J/(g•°C)]

C. Calorimetry

Calorimetry:____

____

Calorimeter:____



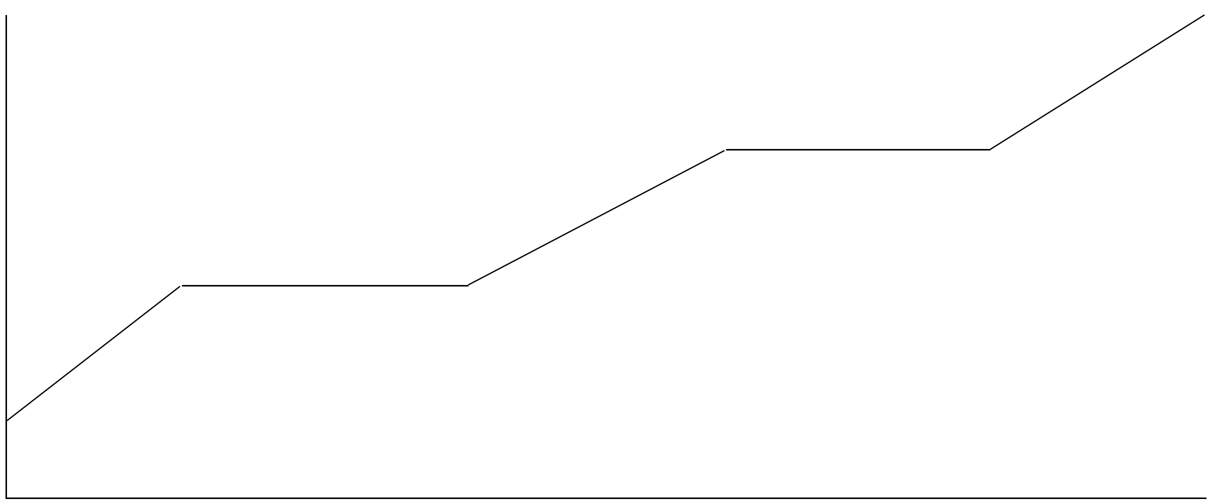
II. Heat Changes Required for Changes of State (Phase)

A. Phase Changes

Solid liquid gas

Heat is absorbed or released when matter changes state.

Interpreting Heat Curve for water



Does water increase in temperature during phase changes?

B. Molar Heats of Phase Changes

Molar heat of fusion: ____

____

____

Molar Heat of Solidification:____

____

____

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

Molar heat of vaporization____

Molar Heat of condensation:____

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

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

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

But they're ALWAYS __________ over __________. 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

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

Molar Heat Calculations Practice

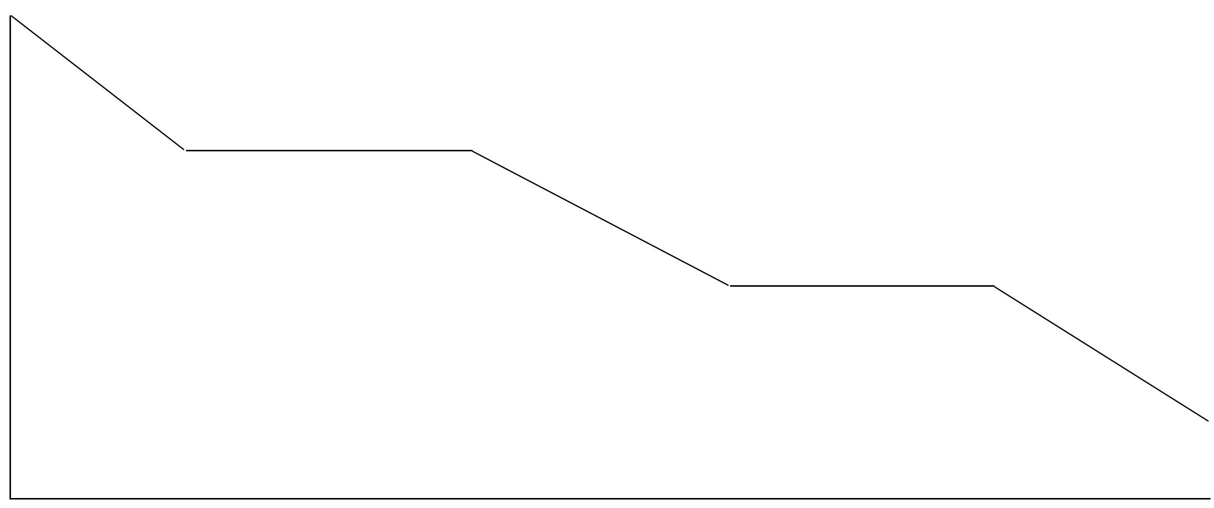
molar heat is given in J/mol or J/g or cal/g, so use it as a conversion factor

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.

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?

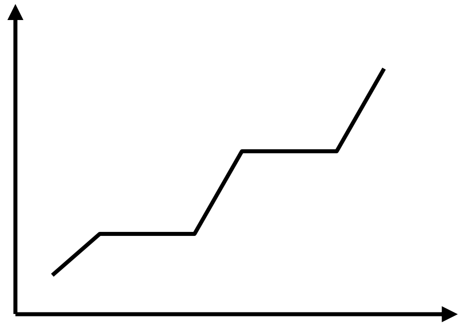
3. How many grams of neon must crystallize (solidify) at it's freezing point to release 560 J of heat, given that the neon's ΔH_{fusion} = 330 J/mol?

Interpreting a Cooling Curve for Water



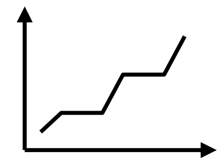
Mixed Molar Heats and Specific Heat Capacity Problems

Use the heating curve to tell which is which

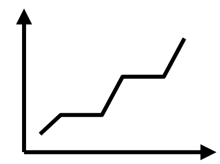


Specific heat capacity $H_2O(s)$ = 2.1 J/(g°C)
Specific heat capacity $H_2O(l)$ = 4.2 J/(g°C)
Heat of fusion H_2O = 6.0 kJ/mol
Heat of vaporization H_2O = 41 kJ/mol

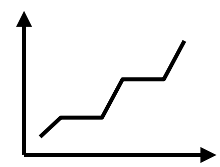
1. How much energy is needed to raise the temperature of 150 grams of ice from -20.0°C to -5.0°C? *(Ans = 4725 J)*



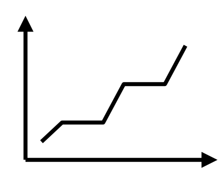
2. How much energy is needed to vaporize 52 grams of water at 100°C?
(Ans = 118 ≈120 kJ).



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***

A. Important Terms

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

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

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 ______________________________________. (*distance* over *time* is *speed*, which is the magnitude of velocity). "_____" is the constant that represents light's speed in a _________________.

Light **frequency** (____, called "_____") × Light **wavelength** (____, called "__________") = c

Max Planck (1900) determines ____

____

Equations:

Energy of light, using frequency:
Energy of light, using wavelength:

$h = 6.626 \times 10^{-34}$ J·s
ν ("nu" not "v") = frequency (unit: $1/s \equiv s^{-1} \equiv Hz$)

____ Small λ = high ν

____ Larger λ = lower ν

Red	Orange	Yellow	Green	Blue	Indigo	Violet

Wavelength
= 7.0×10^{-7} m
Low Energy

wavelength
4×10^{-7} m
High Energy

Louis de Broglie (1924) determines ____

____

____

The de Broglie Equation

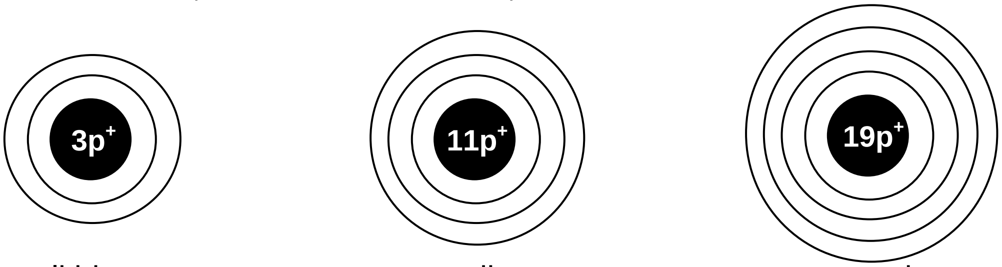
& Interpretation:

What did the spectrum tube demonstration show?

Bohr's Model of the Atom (Powerpoint)

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

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



lithium sodium potassium

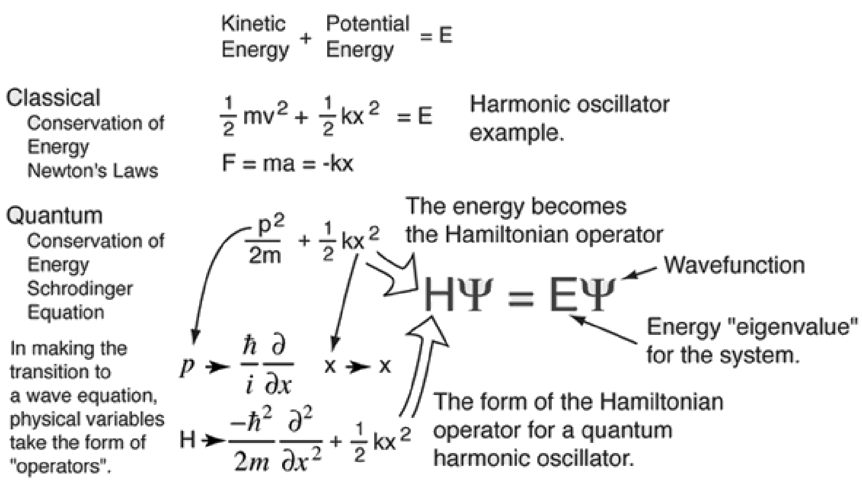
The number of electrons in the outer principal energy level (or _______________ shell) is the same within a group.

The number of principle energy levels is the same as the _________________________.

So why isn't this model good enough?

The Quantum Mechanical Model

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



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

____

____

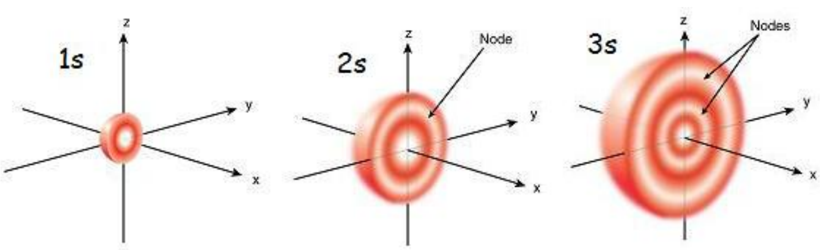
____

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

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

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

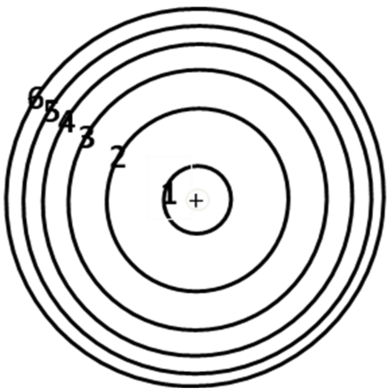
C. The areas where an electron can *probably* be found are called _________________. (The areas where electrons will be unlikely are called __________)

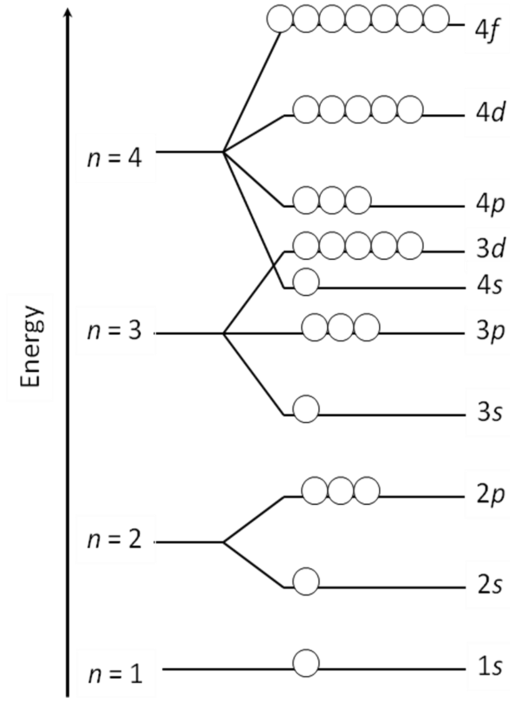


D. Each orbital has a specific shape and can hold up to ___________________ (spinning in ***opposite*** directions).

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

E. Organization of Electrons

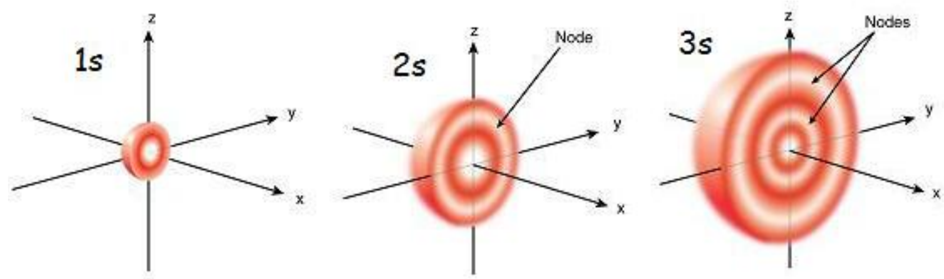
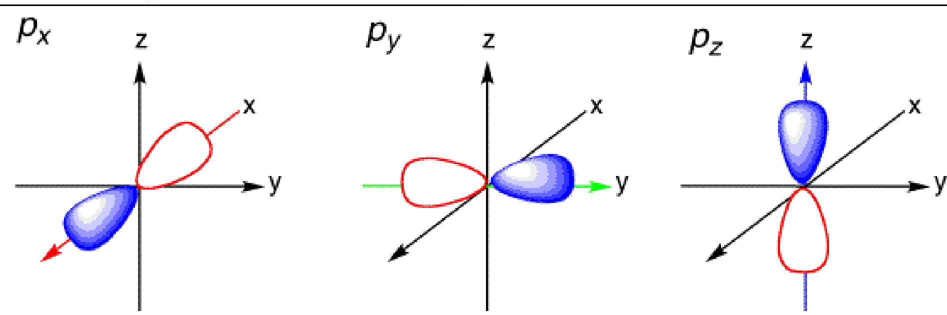


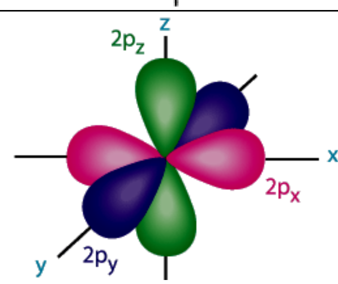


Description of Sublevels

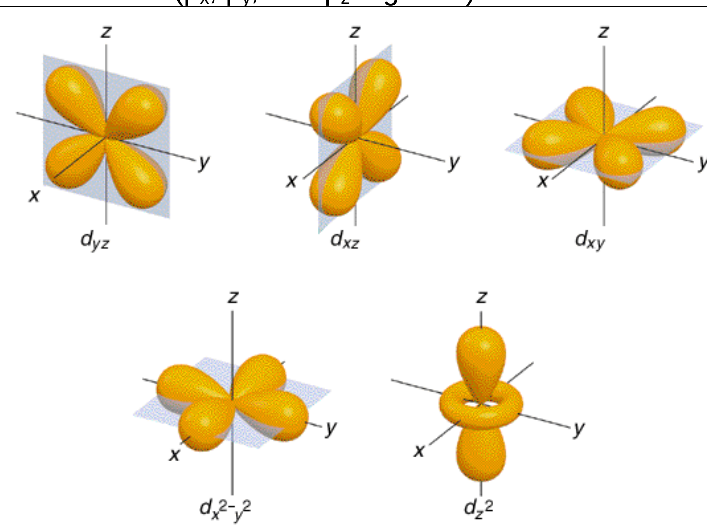
1. “s” sublevels have 1 orbital – it is ___________________ shaped
2. “p” sublevels have 3 orbitals – they are ___________________ shaped
3. “d” sublevels have 5 orbitals – 4 are ___________________ shaped, one “pacifier”
4. “f” sublevels have 7 orbitals – they are ___________________ shaped

<table>
<tr><td>What orbital type is this?

How many electrons can go in this orbital?</td><td>
</td></tr>
<tr><td>How many electrons can fill each p- orbital?</td><td>

The three p orbitals are aligned along perpendicular axes</td></tr>
<tr><td>How many electrons can fill this entire p-energy sublevel?</td><td>

The p-energy sublevel is made of all three 3D orientations (p_x, p_y, and p_z together)</td></tr>
<tr><td>How many electrons can fill each d-orbital?

How many electrons can fill this d-energy sublevel?</td><td>
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F. Filling in the Orbitals in Quantum Mechanics

1. **Aufbau Principle**:

Electrons fill the ____ first.

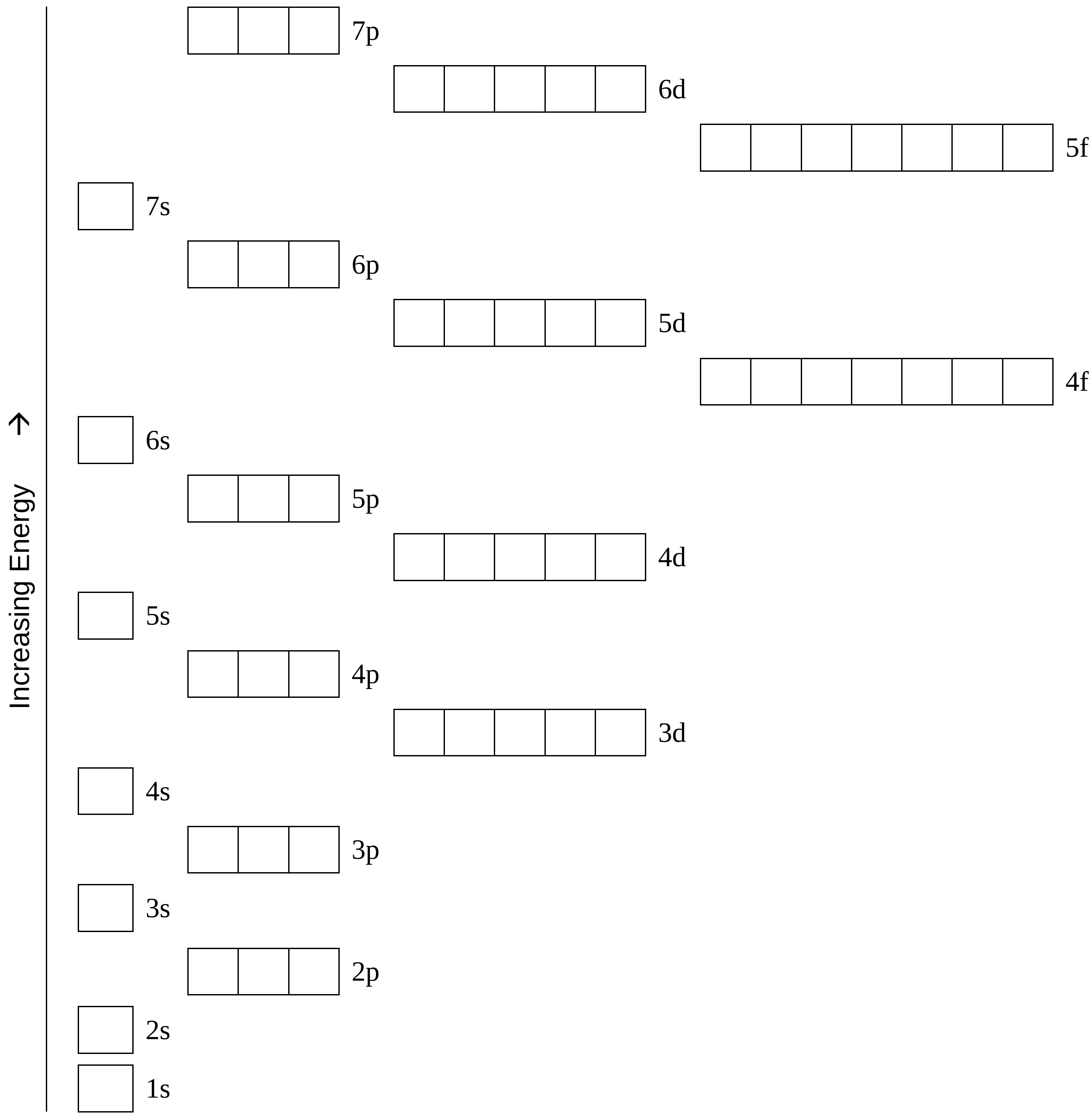
2. **Pauli Exclusion Principle** (PEP):

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

3. **Hund's Rule**:

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

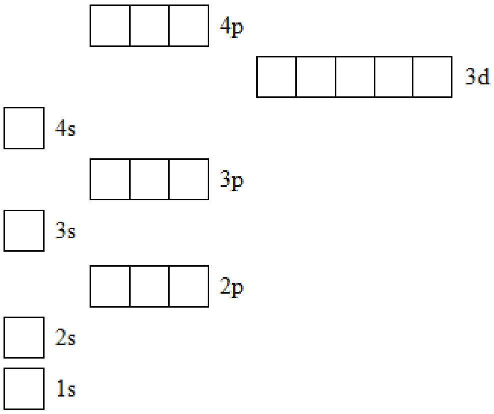
Electron Configurations & the Aufbau Diagram



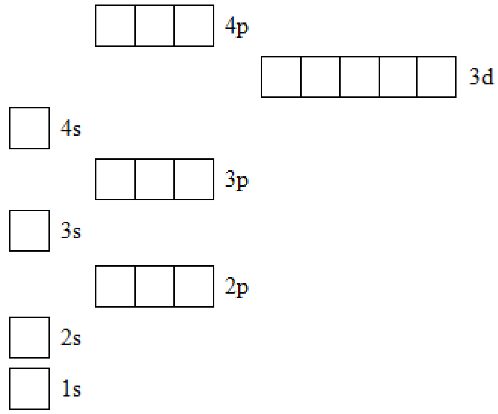
Rules to fill it in:

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]*

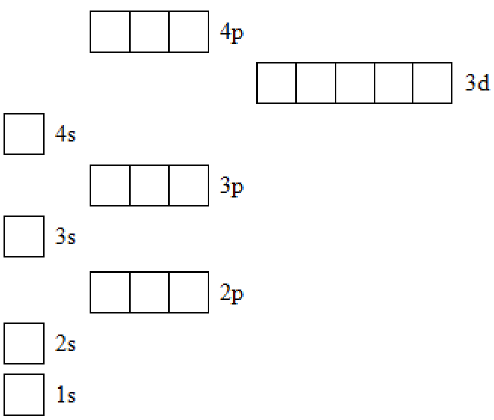
Li #electrons =



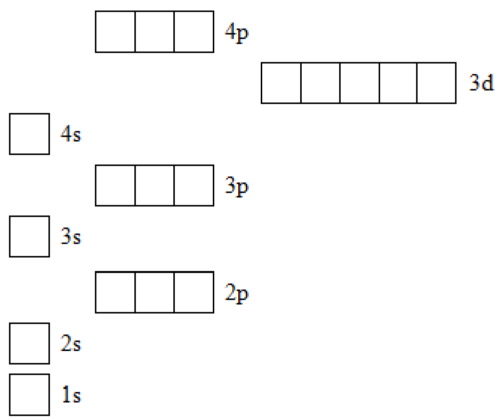
B # electrons =



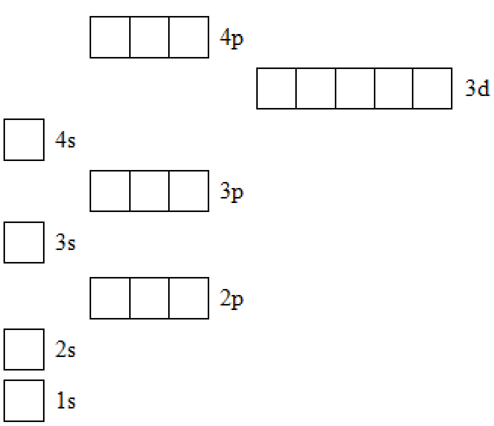
N # electrons =



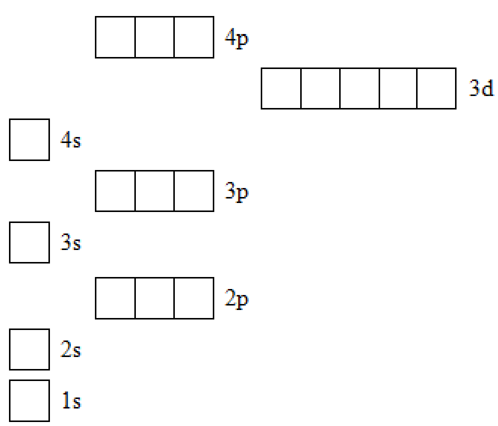
Ne # electrons =



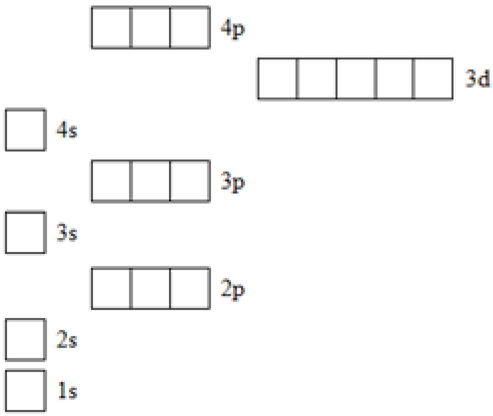
P # electrons =



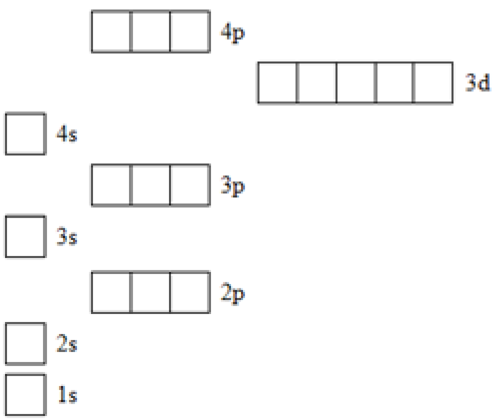
Ca #electrons =



$\boxed{As}$ # electrons =



$\boxed{Cr}$ # 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

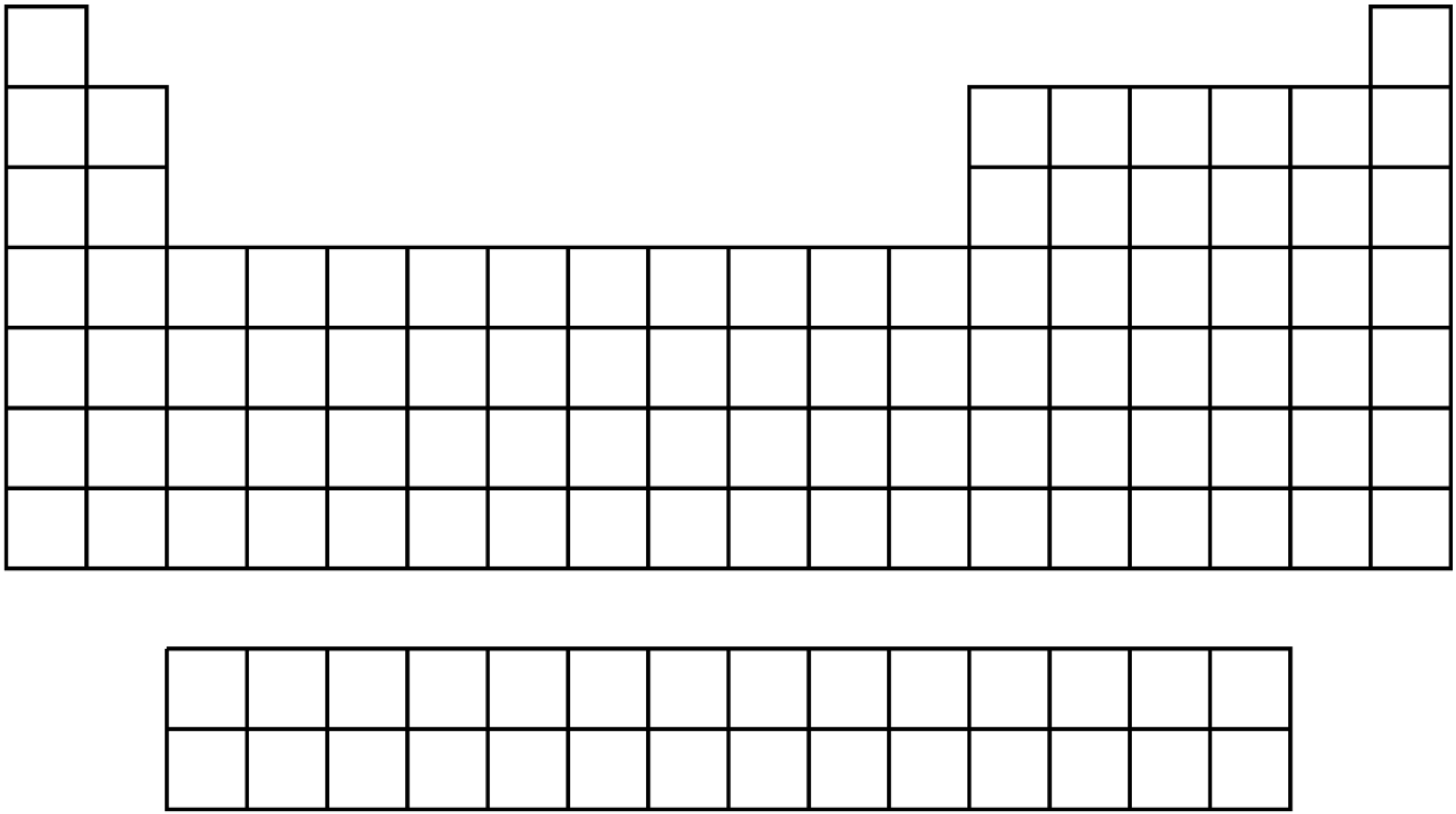
Si: ______ Mg:______ Fe:_____

Chapter 13: Periodic Table and Electron Configuration

1	H																	He
group																		
valence e- config.																		
2	Li	Be											B	C	N	O	F	Ne
3																		
4			Sc									Zn						
5																		
6													Tl					Rn
7	Fr	Ra	Lr									Uub						
# valence electrons																		

La													Yb
Ac													No

Representative elements: ____



Chapter 15: Ionic Bonds

I. **Valence** Electrons

A. definition: ____

____

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

Lewis Structures show bonded atoms as __________.

Lewis Dot Structure Example (single atom):

Lewis Structure of Molecule using Dots Only:

Lewis Structure of Molecule using Lines for Bonds:

Group	1	2	13	14	15	16	17	18
Example								

II. **Octet** Rule

A. definition: ____

____

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

- configuration example: Na =

  Na^+ =

- configuration example: S =

  S^{2-} =

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

Helium is happy with __________ valence electrons, so we call this exception the

___________ rule. Atoms with atomic numbers close to He (such as ____, ____, and

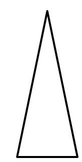
____) will be happy with 2 electrons. They can't fit 8!

H^+: = H^- = Li^+ = Be^{2+} =

____

____

Terminology: **iso**- means ______________ or __________ (think: isosceles triangle)
-**electronic** refers to the number of ___________________.



∴ , what does isoelectronic mean?

Concept Check:

- Are He and Ne isoelectronic with each other? _________
- Are O^{2-} and Ne isoelectronic with each other? _________
- Are F^- and Cl^- isoelectronic with each other? __________
- Are Cl^- and S^{2-} isoelectronic with each other? __________
- Which noble gas will iodine become isoelectronic to when an iodine atom is ionized?

  ____
- Na+ will lose one electron, to become isoelectronic with ___________________.
- Which alkaline metal is most likely to ionize to become isoelectronic with the noble gas Krypton? _______________________
- Are Mg^{2+} and N^{3-} isoelectronic? ________

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?*

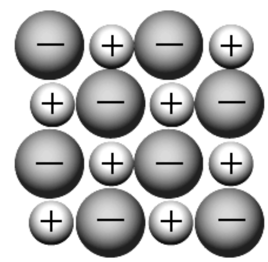
____

____

____

____

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



IV. Properties of Ionic Compounds

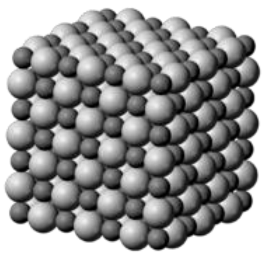
Ionic compounds are held together by ________________________________

Electrostatic attraction:____

____

A. crystal structure: ____

____



NaCl

B. electrolytes: ____

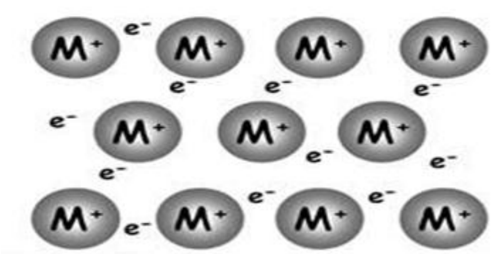
____

C. high melting points

V. Metallic Bonds

A. caused by attraction of ________________________________ for the positively charged ______________________________________

B. metals are good ____________________ 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 _____________________. An **alloy** is a solid mixture of metals.

Two (or more) metals are ____________________, then mixed together while they're

still ___________________. After the hot liquid metal mixture cools, you have an alloy.

<u>Brass</u> is made of ________________ and _________________.
It's great for musical instruments due to how sound waves resonate (propagate) through the metal atoms!
<u>Bronze</u> is made of ____________________ , tin, and other metals.

<u>Steel</u> is made of ______________________, carbon, and other elements.

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

Investigation Questions:

Why is it not a good idea to have jewelry that is <u>pure</u> gold?

What makes <u>stainless steel</u> special? And why doesn't it stain easily?

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 ______________________________

$Sr_3(PO_4)_2$ ______________________________

$Cu(NO_3)_2$ ______________________________