

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. $\text{Percent Error} = 100 \times \frac{ \text{accepted value} - \text{exper. value} }{\text{accepted value}}$	
Atomic Structure and Periodic Relationships 2.3 SOL 2d, 2g, 2i	<p>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.</p> <p>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.</p> <p>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</p> <p>Lewis dot diagrams show the valence electrons of an atom. The electrons (dots) are arranged around the element's symbol.</p> <p>Metallic bonds consist of the attraction of free-floating valence electrons for the positively charged metal ions.</p>	<p>Ch 13: Read pp 361-364 on development of modern quantum mechanical model.</p> <p>Ch 13: Read pp 364-366 on orbital shapes.</p> <p>Ch 13: Read pp 367-369 on electron configurations and Aufbau principle.</p> <p>Ch 13: Read pp 372-381 on the relationship between atomic orbitals and atomic emission spectra.</p> <p>Ch 15: Read pp 413-424 on ionic compounds. Read pp 427 on metallic bonds.</p>
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.	<p>Ch 6: Review pp 146-156 for naming and writing formulas for ionic compounds.</p>
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: $\text{NaCl}(s) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)$ Ionic compounds that dissociate completely in water are strong electrolytes.	<p>Ch 17: read pp 483-485</p>
Phases of Matter and Kinetic Molecular Theory 5.3 SOL 5d,5e,5f	<p>Specific Heat Capacity (C) is a physical property of a substance. $Q = mC\Delta T$ is use to calculate heat, mass, specific heat or temperature change respectively. Specific heat can be used to identify a substance.</p> <p>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.</p> <p>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.</p> <p>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.</p>	<p>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.</p>

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} , $\Delta H_{\text{vaporization}}$, 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: _____

Potential Energy: _____

Kinetic Energy: _____

Heat: _____

Thermochemistry: _____

B. Exothermic and Endothermic Processes

Processes that absorb (_____) heat or release (_____) heat.

Definitions

Consider $\text{NH}_4\text{Cl(s)} \xrightarrow{\text{water}} \text{NH}_4^+(\text{aq}) + \text{Cl}^-(\text{aq})$

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: _____

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

Specific Heat Capacity: _____

Symbol:

Units:

Equation:

q = _____

C = _____

m = _____

$\Delta T =$ _____

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 $\text{J}/(\text{g}\cdot^\circ\text{C})$; “ ΔT ” is change in temperature in $^\circ\text{C}$ (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.

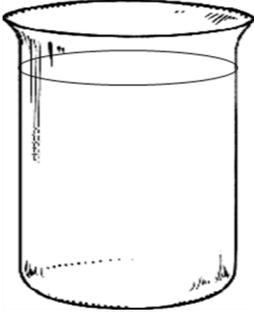
Sample Problems

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 $0.46 \text{ J}/(\text{g}\cdot^\circ\text{C})$. 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 $4.18 \text{ J}/(\text{g}\cdot^\circ\text{C})$ 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 $0.90 \text{ J}/(\text{g}\cdot^\circ\text{C})$, 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. Challenger: 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? [$C_{\text{Ag}} = 0.24 \text{ J}/(\text{g}\cdot^\circ\text{C})$] (*Ans* = 21.4°C)

C. Calorimetry

Calorimetry: _____

Calorimeter: _____

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

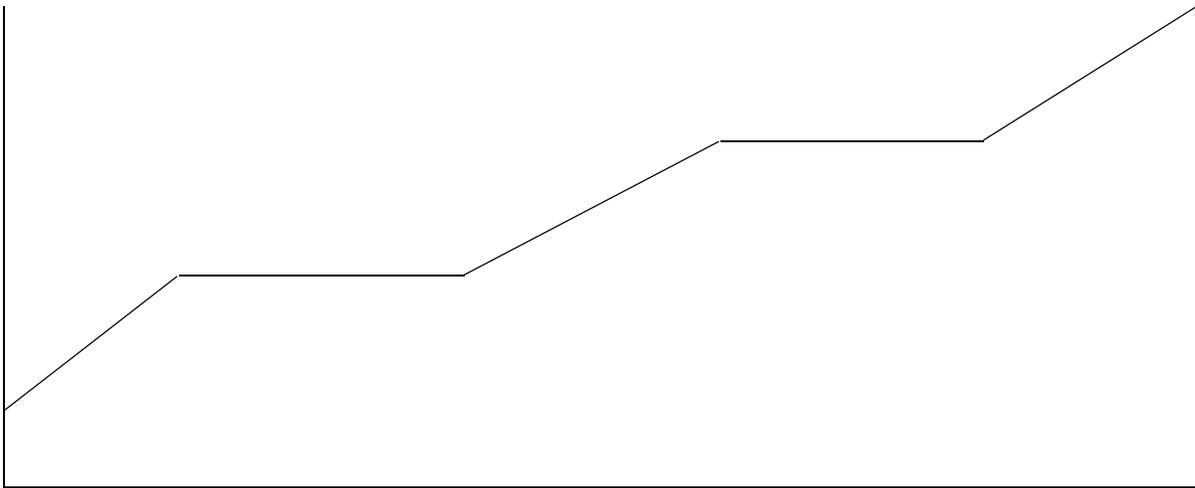
Solid

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: _____

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

Molar heat of vaporization _____

Molar Heat of condensation: _____

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

Molar Heats apply to phase changes The units may include:

J/mol

J/gram

calories/gram

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? $\Delta H_{\text{fusion}} = 6.0 \text{ kJ/mol}$

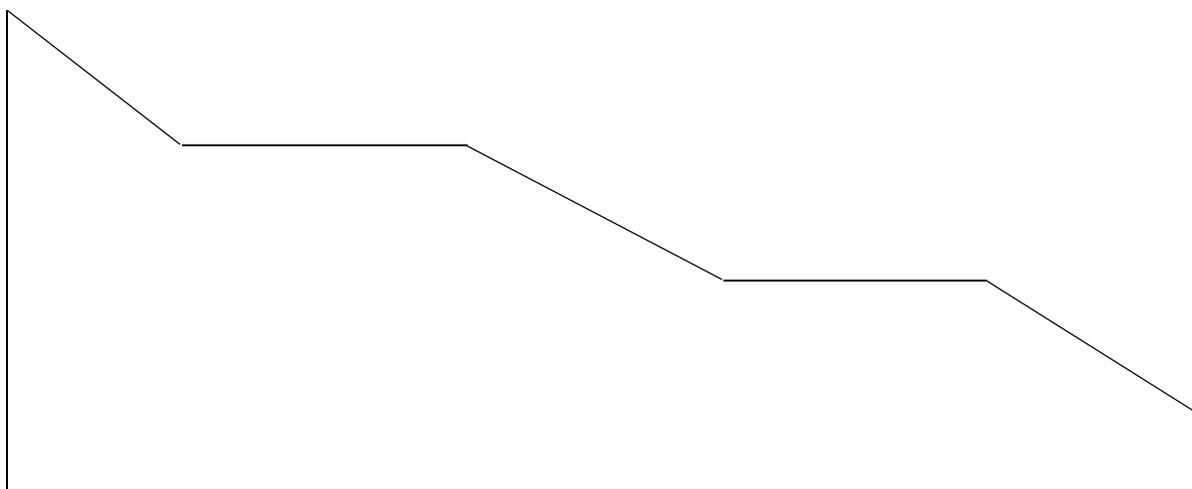
Example 2: How many kilojoules of heat would be released when 36.04 grams of steam condenses to water at 100°C?



Molar Heat Calculations Practice

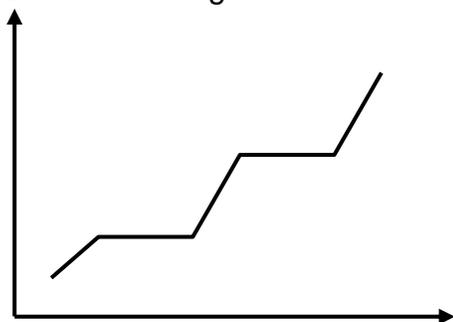
molar heat is given in J/mol or J/g or cal/g, so use it as a conversion factor

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.
2. How much heat is required to vaporize 13.1 grams of methane (CH₄) at its boiling point, which has a heat of vaporization of 8.2 kJ/mol?
3. How many grams of neon must crystallize (solidify) at its freezing point to release 560 J of heat, given that the neon's $\Delta H_{\text{fusion}} = 330 \text{ J/mol}$?

Interpreting a Cooling Curve for Water

Mixed Molar Heats and Specific Heat Capacity Problems

Use the heating curve to tell which is which



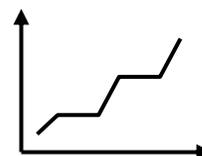
Specific heat capacity $\text{H}_2\text{O}(\text{s}) = 2.1 \text{ J}/(\text{g}^\circ\text{C})$

Specific heat capacity $\text{H}_2\text{O}(\text{l}) = 4.2 \text{ J}/(\text{g}^\circ\text{C})$

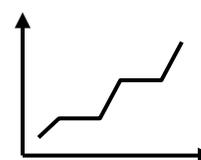
Heat of fusion $\text{H}_2\text{O} = 6.0 \text{ kJ/mol}$

Heat of vaporization $\text{H}_2\text{O} = 41 \text{ kJ/mol}$

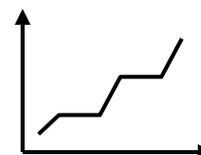
1. How much energy is needed to raise the temperature of 150 grams of ice from -20.0°C to -5.0°C ? (*Ans = 4725 J*)



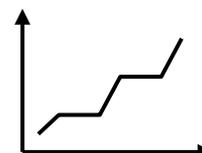
2. How much energy is needed to vaporize 52 grams of water at 100°C ? (*Ans = 118 \approx 120 kJ*).



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)

<i>What</i>	<i>How</i>	<i>Model</i>

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. $^{13}_6\text{C}$ — 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 “_____”) \times Light **wavelength** (_____, called “_____”) = c

Max Planck (1900) determines _____

Equations:

Energy of light, using frequency:

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

_____ Small $\lambda =$
high ν

Energy of light, using wavelength:

_____ Larger $\lambda =$
lower ν

Red	Orange	Yellow	Green	Blue	Indigo	Violet
-----	--------	--------	-------	------	--------	--------

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

wavelength
 $4 \times 10^{-7} \text{ m}$
 High Energy

Louis de Broglie (1924) determines _____

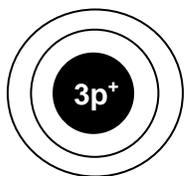
The de Broglie Equation & Interpretation:

What did the spectrum tube demonstration show?

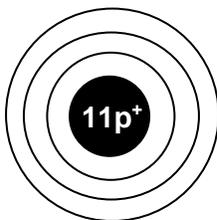
Bohr's Model of the Atom (Powerpoint)

<i>What</i>	<i>How</i>	<i>Model</i>

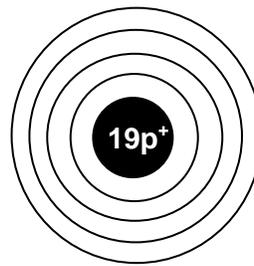
A Few Bohr Models (let's add the electrons)



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.

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

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

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

$H \rightarrow \frac{-\hbar^2}{2m} \frac{\partial^2}{\partial x^2} + \frac{1}{2}kx^2$

The energy becomes the Hamiltonian operator

$H\Psi = E\Psi$

Wavefunction

Energy "eigenvalue" for the system.

The form of the Hamiltonian operator for a quantum harmonic oscillator.



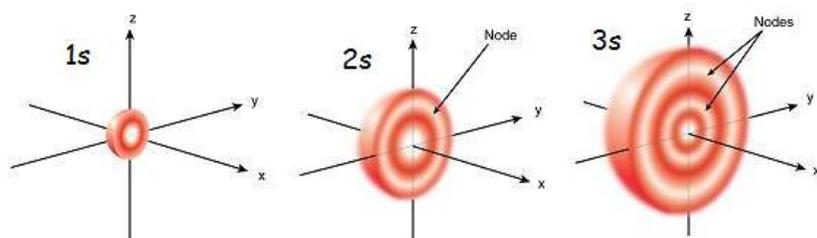
B. Werner Heisenberg adds the **Heisenberg Uncertainty Principle**:

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

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

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

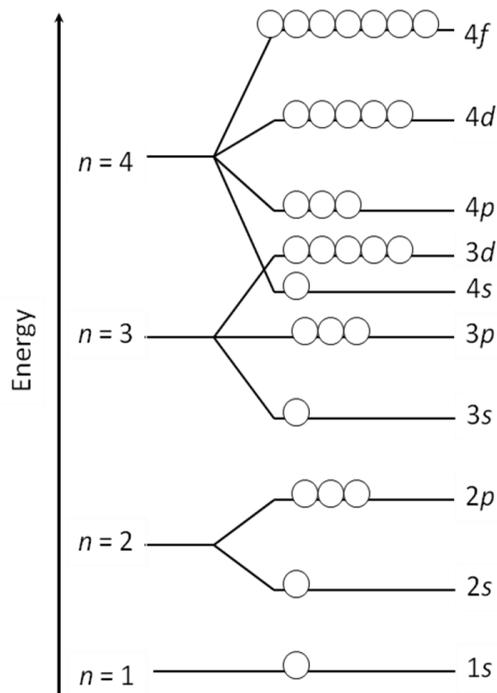
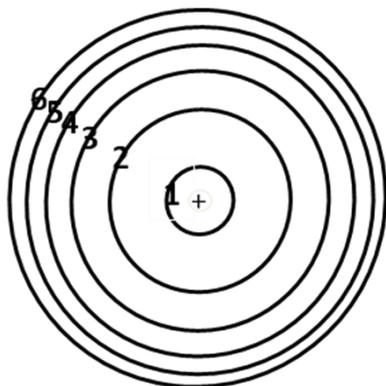
C. The areas where an electron can *probably* be found are called _____. (The areas where electrons will be unlikely are called _____)



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

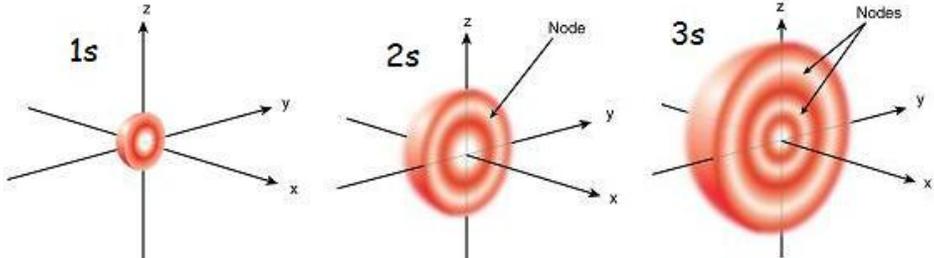
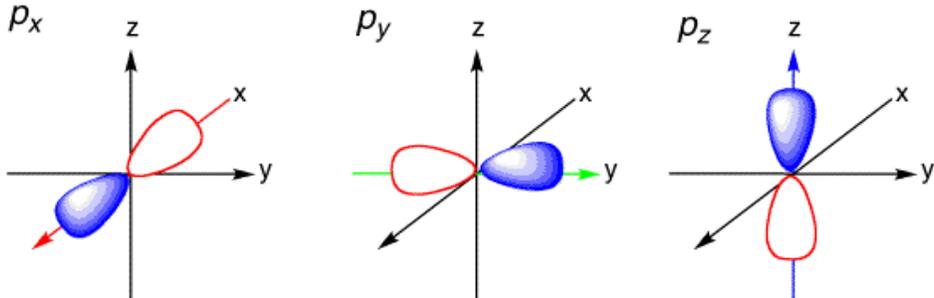
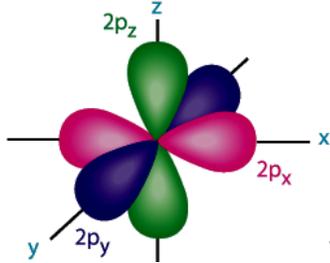
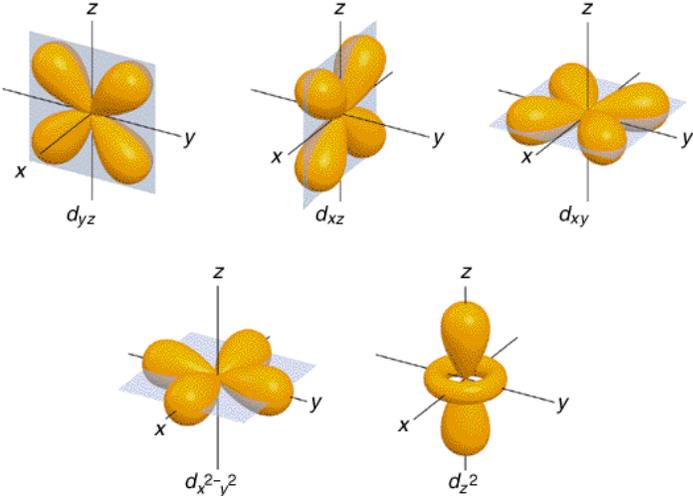
a. How is the spin of an electron noted in models? _____

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

<p>What orbital type is this?</p> <p>How many electrons can go in this orbital?</p>	
<p>How many electrons can fill each p- orbital?</p>	 <p>The three p orbitals are aligned along perpendicular axes</p>
<p>How many electrons can fill this entire p-energy sublevel?</p>	 <p>The p-energy sublevel is made of all three 3D orientations (p_x, p_y, and p_z together)</p>
<p>How many electrons can fill each d-orbital?</p> <p>How many electrons can fill this d-energy sublevel?</p>	

F. Filling in the Orbitals in Quantum Mechanics

1. Aufbau Principle:

Electrons fill the _____ first.

2. Pauli Exclusion Principle (PEP):

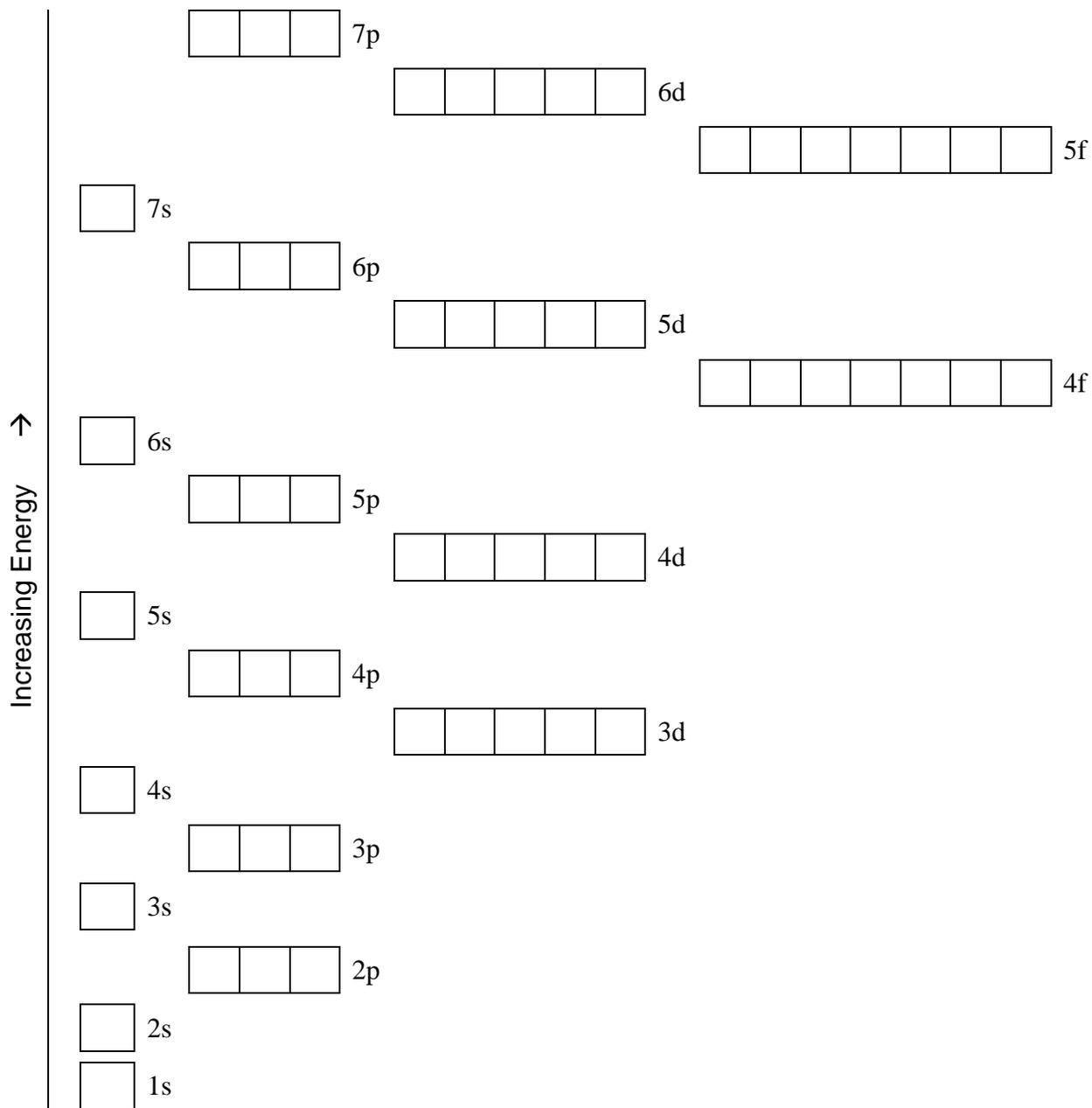
Only _____ electrons can be in each orbital (two per box or line!) and they must have _____ magnetic _____. (two different arrow directions!)



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...

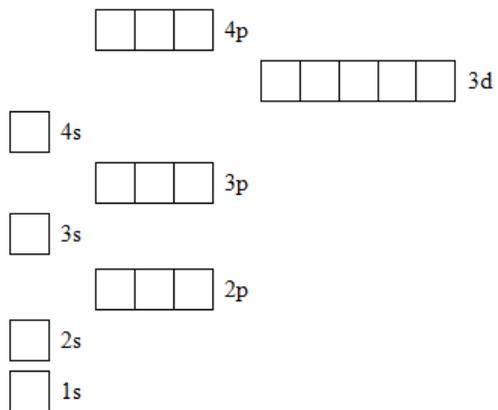
Electron Configurations & the Aufbau Diagram



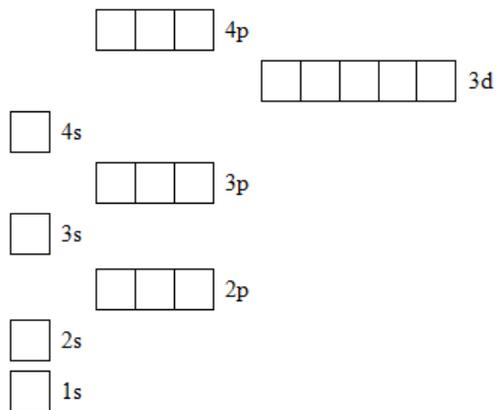
Rules to fill it in:

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*]

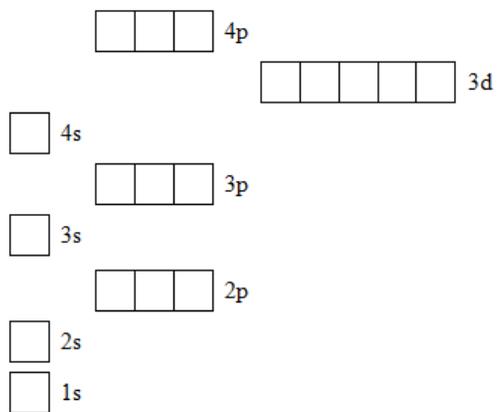
Li #electrons =



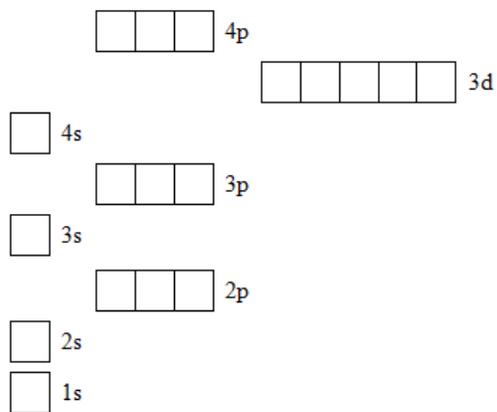
B # electrons =



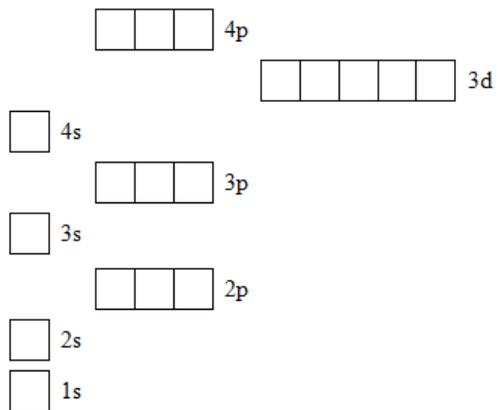
N # electrons =



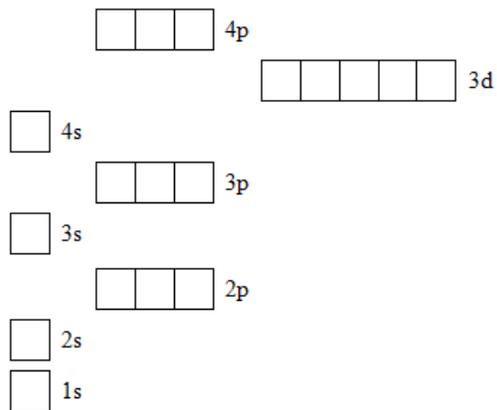
Ne # electrons =



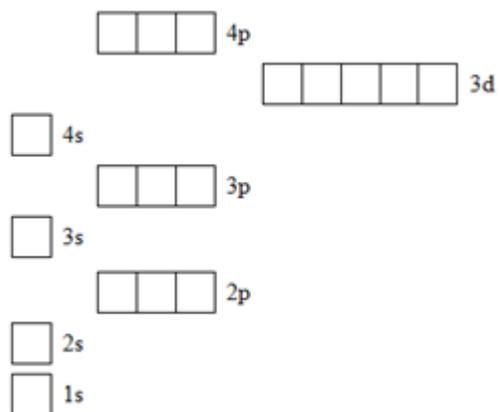
P # electrons =



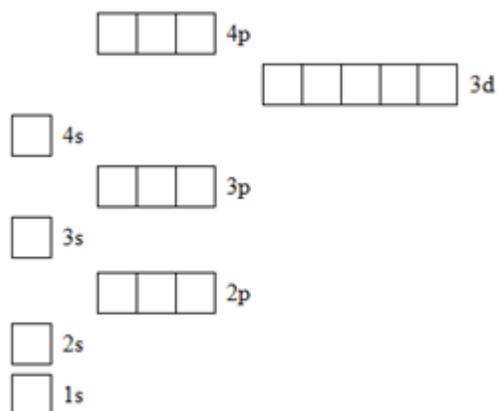
Ca #electrons =



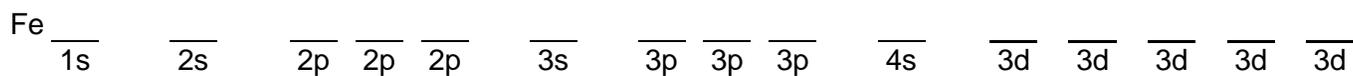
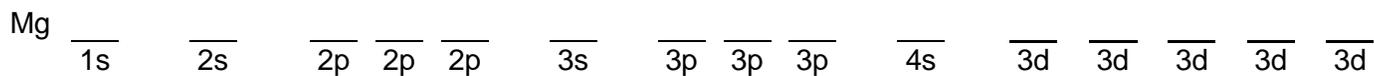
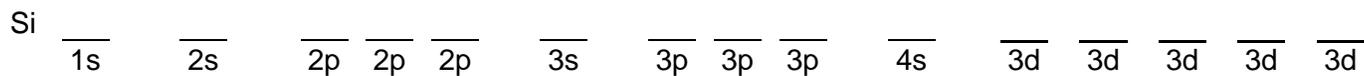
As # electrons =



Cr # electrons =



Complete the alternate form of the Aufbau diagrams below:



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														
7	Fr	Ra	Lr							Uu b				
# 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 =



- configuration example: S =



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⁺ = Be²⁺ =

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 O²⁻ and Ne isoelectronic with each other? _____
- Are F⁻ and Cl⁻ isoelectronic with each other? _____
- Are Cl⁻ and S²⁻ 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 Mg²⁺ and N³⁻ isoelectronic? _____

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

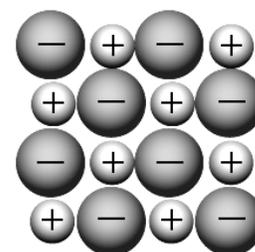
Boron (B) *actually* prefers to have _____ valence electrons (and it's stable that way!), rather than 8 like many others.

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, CaF₂, MgO, Li₃P, K₂S)

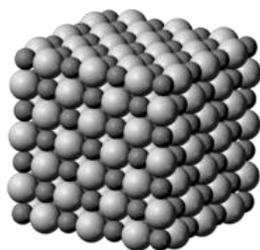


IV. Properties of Ionic Compounds

Ionic compounds are held together by _____

Electrostatic attraction: _____

A. crystal structure: _____



NaCl

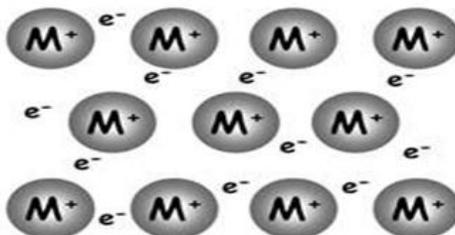
B. electrolytes: _____

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.

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

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

Steel is made of _____, 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.
_____ bonds keep it together, of course.

Investigation Questions:

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

What makes stainless steel 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

$\text{Fe}(\text{NO}_2)_3$

Ba_3P_2

$\text{V}(\text{OH})_5$

Cr_2O_3

$\text{Sr}_3(\text{PO}_4)_2$

$\text{Cu}(\text{NO}_3)_2$
