Atoms

You'll Need to Know

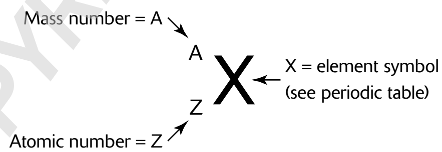
- ❑ what is the Aufbau Principle
- ❑ how to determine noble gas core abbreviation
- ❑ what is the Bohr model
- ❑ what are quantum numbers
- ❑ diamagnetism vs. paramagnetism
- ❑ important people, experiments, and theories
- ❑ energy and other relationships (amplitude, frequency, momentum, photoelectric effect, Planck's constant, Rydberg constant, velocity, wavelength, etc.)

The Aufbau Principle

Electrons are added to atoms in order of increasing energy: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14}$.

Noble Gas Core Abbreviation

Select the noble gas that most nearly precedes the element being considered. Place its symbol in square brackets, and continue from that point with the electron configuration. Example: Sr would be $[Kr]5s^2$.



For neutral atoms, the number of electrons = the number of protons. The number of neutrons = A – Z.

Examples

Q. Write the electron configuration for Cl.

A. $1s^2 2s^2 2p^6 3s^2 3p^5$

Q. Write the electron configuration for Cl^-.

A. The "–" charge means to add an extra electron. $1s^2 2s^2 2p^6 3s^2 3p^6$. Cl^- is isoelectronic with Ar.

Q. Write the electron configuration for Ca^{2+}.

A. The 2+ charge means to subtract two electrons. It would be $1s^2 2s^2 2p^6 3s^2 3p^6$. Ca^{2+} is isoelectronic with Ar and Cl^-.

Q. Write the electron configuration for Cr.

A. [Ar] $4s^1 3d^5$ Extra stability is gained by having the *d* orbitals half-filled. You will find this same phenomenon with Cu: $[Ar]4s^1 3d^{10}$ in which the *d* orbitals are fully filled, providing a more stable configuration.

Q. Write the electron configuration for Mn^{3+}.

A. $[Ar]3d^4$ Transition metals characteristically have incompletely filled *d* orbitals and readily give rise to ions with incompletely filled *d* orbitals. This gives them distinct colors, ability to form paramagnetic compounds, catalytic activity, and the ability to form complex ions.

Q. How many unpaired electrons are there in the tungsten atom (W)?

A. Tungsten has 74 electrons. $[Xe]6s^2 4f^{14} 5d^4$. Each of the 5*d* electrons would be in a separate orbital, resulting in 4 unpaired electrons.

Bohr Model

A "solar system" model of the atom in which: (1) electrons moved in circular orbits around the nucleus; (2) only certain orbits with distinct radii were allowed (beginning of quantum theory); (3) energy of electron remained constant as long as it was within an orbit; (4) energy of the electron increased the farther its orbit was from the nucleus; (5) when energy was absorbed by the atom, the electron jumped to a higher energy orbit, farther from the nucleus; and (6) when energy was released by an atom, the electron jumped to a lower energy orbit, closer to the nucleus. Explained the source and observed wavelengths of lines in the hydrogen spectrum.

Quantum Numbers

Schrödinger described an atomic model with electrons in three dimensions and required three coordinates, or three quantum numbers, to describe where electrons could be found. The fourth quantum number(s) describes the spin of the electron. The three postulates of the quantum theory are: (1) atoms and molecules can exist only in discrete states, characterized by definite amounts of energy. When an atom or molecule changes state, it absorbs or emits just enough energy to bring it to another state; (2) when atoms or molecules absorb or emit light in moving from one energy state to another, the wavelength of the light is related to the energies of the two states as follows:

$$E_{\text{high}} - E_{\text{low}} = \frac{h \cdot c}{\lambda}$$

and (3) the allowed energy states of atoms and molecules are described by quantum numbers.

***n*: principal (shell) quantum number:** describes the energy level within the atom.

- 1 – 7 energy levels. As "*n*" increases, the shell is further away from the nucleus. Electrons have more energy the farther they are away from the nucleus.
- Maximum number of electrons in a shell (energy level) is $2n^2$.

***l*: momentum (subshell) quantum number:** describes the shape of the orbital. An orbital can hold a maximum of 2 electrons.

- Describes the sublevel in "*n*".
- Each energy level has "*n*" sublevels.
- Sublevels of different energy levels may have overlapping energies.
- Sublevels in the atoms of the known elements are:

  $l = 0$ (*s* orbital- spherical)

  $l = 1$ (*p* orbital- dumbbell shape)

  $l = 2$ (*d* orbital- cloverleaf shape)

  $l = 3$ (*f* orbital)

m_l: magnetic quantum number: describes the direction or orientation in space for the orbital.

- *s* only has 1 orbital, coded as "0".
- *p* has 3 possible orbitals: p_x, p_y, and p_z; coded as –1, 0, 1 respectively.
- *d* has 5 possible orbitals: d_{z^2}, d_{xz}, d_{yz}, d_{xy}, $d_{x^2-y^2}$ coded as –2, –1, 0, 1, 2 respectively.
- *f* has 7 possible orbitals coded as –3, –2, –1, 0, 1, 2, 3.

m_s: spin quantum number: describes the spin of the electron.

- Electrons in the same orbital must have opposite spins.
- Possible spins are clockwise (+½) or counterclockwise (–½).

Examples

Q. **What are the possible values for an electron in a 4*d* orbital?**

A. Principal energy level, $n = 4$. Since it is a *d* orbital, $l = 2$. m_l values can vary from $-l$ to l; therefore, m_l can be –2, –1, 0, 1, or 2.

Q. **An atom has a principal quantum number of 3. How many orbitals are associated with $n = 3$?**

A. For $n = 3$, there is one 3*s* orbital, three 3*p* orbitals; and five 3*d* orbitals for a total of 9.

Diamagnetism vs. Paramagnetism

Diamagnetic: all subshells are filled with electrons.

(Examples: He: $1s^2$; Be: $1s^2 2s^2$; Ne: $1s^2 2s^2 2p^6$). Elements are NOT affected by magnetic fields.

Paramagnetic: subshells are NOT completely filled with electrons.

(Examples: Li: $1s^2 2s^1$; N: $1s^2 2s^2 2p^3$; F: $1s^2 2s^2 2p^5$). Elements ARE affected by magnetic fields.

Example

Q. **Is palladium (Pd) diamagnetic or paramagnetic?**

A. Palladium has 46 electrons and an electron configuration $[Kr]4d^{10}$. Since the *d* orbitals are completely filled, palladium would be diamagnetic.

Important People, Experiments, and Theories

Neils **Bohr**	Electrons orbit the nucleus at specific fixed radii similar to planets orbiting the sun. Electrons can jump to higher energy levels after absorbing specific amounts of energy. Likewise, electrons dropping from higher to lower energy levels will release certain amounts of energy in the form of photons. Model was replaced with quantum mechanics.
Louis **de Broglie**	All matter has both particle and wave characteristics.
John **Dalton**	Elements are made up of unique atoms. Elements combine to make compounds. Compounds have constant ratios of atoms. Atoms are neither created nor destroyed in chemical reactions.
Albert **Einstein**	Photoelectric effect: When light with certain frequencies strikes a piece of metal, it emits electrons from the metal. Radiant energy behaves as a stream of tiny packets of energy (photons). Photons have properties of waves (wave-particle duality of nature).

Heisenberg Uncertainty Principle	It is not possible to know both the position and momentum of an electron at a particular moment. Electron orbitals are described in terms of probability.
Hund's Rule	Electrons will enter empty orbitals of equal energy when they are available.
James **Maxwell**	Maxwell provided a mathematical description of the general behavior of light. He described how energy in the form of radiation can travel through space as electric and magnetic fields.
Robert **Millikan**	Millikan calculated the charge of an electron.
Pauli Exclusion Principle	No two electrons in an atom have the same set of four quantum numbers.
Max **Planck**	Planck is the "father" of quantum mechanics. Energy can only be emitted or absorbed from atoms in fixed amounts (quanta).
Ernest **Rutherford**	Gold foil experiment: He shot a beam of alpha particles at a thin sheet of gold and found that the atoms in the foil must contain an extremely dense, positively charged core, sufficient to deflect the positively charged alpha particles. Atoms are mostly space. The nucleus is very dense positively charged center of atom. Electrons are small and travel around nucleus.
Erin **Schrödinger**	Through development of the Schrödinger equations, he was able to apply probability to describing the volume of space of where an electron would be located. Wave mechanics became the foundation for the development of the quantum model of the atom.
J.J. **Thomson**	"Plum Pudding Model": The atom consisted of a positively charged, spherical mass with negatively charged electrons scattered throughout.

Energy and Other Relationships

Amplitude (Ψ)	The vertical distance from the midline of a wave to the peak or trough.
Frequency (ν)	Number of waves that pass through a particular point in one second. Common units are hertz (Hz or sec^{-1}).
de Broglie equation	$\lambda = \frac{h}{m \cdot u}$
Electron charge (e^-)	$e^- = -1.602 \times 10^{-19}$ coulomb
Energy of an electron	$E_n = \frac{-2.178 \times 10^{-18}}{n^2}$ joules where E_n = energy of the electron and n = principal quantum number
Momentum	$p = m \cdot u$
Photoelectric effect	The emission of electrons from the surface of a metal when light shines on it. Electrons are emitted, however, only when the frequency of that light is greater than a certain threshold value characteristic of the particular metal.
Planck's constant (h)	$h = 6.63 \times 10^{-34}$ J · s
Rydberg constant	$R_H = 2.18 \times 10^{-18}$ J $\Delta E = h \cdot v = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$
Speed of electromagnetic radiation (c)	$c = 3.00 \times 10^8$ m · s^{-1} ($c = \lambda \cdot \nu$)
Velocity (υ)	The speed of a wave. Common units are cm · sec^{-1}. $\upsilon = \lambda \cdot \nu$
Wavelength (λ)	Distance between identical points on successive waves. Common units are nanometers (10^{-9} m) or Angstrom (Ä) (10^{-10} m).

Examples

Q. Calculate the speed of a wave whose wavelength is 17.5 cm and whose frequency is 89.0 Hz.

A. $\upsilon = \lambda \cdot \nu = 17.5 \text{ cm} \cdot 89.0 \text{ s}^{-1} = 1.56 \cdot 10^3 \text{ cm} \cdot \text{s}^{-1}$

Q. Green light has a wavelength of 520 nm. What is the frequency of this light?

A. $\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \cancel{\text{m}}/\text{s}}{520 \cancel{\text{nm}}} \times \frac{10^9 \cancel{\text{nm}}}{1 \cancel{\text{m}}} = 5.77 \times 10^{14} \cdot \text{s}^{-1}$

Q. Calculate the energy (in joules) of a photon whose wavelength is 6.00 × 10⁴ nm.

A. $E = \frac{h \cdot c}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \cancel{\text{m}}/\text{s})}{6.00 \times 10^4 \cancel{\text{nm}}} \times \frac{10^9 \cancel{\text{nm}}}{1 \cancel{\text{m}}} = 3.32 \times 10^{-21} \text{ J}$

Q. What is the wavelength (in nm) of a photon emitted during a transition from n_i = 5 state to the n_f = 3 state in the hydrogen atom?

$$\Delta E = h \cdot \nu = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = 2.18 \times 10^{-18} \text{ J} \left(\frac{1}{5^2} - \frac{1}{3^2} \right)$$

A. $\Delta E = -1.53 \times 10^{-19} \text{ J}$ (*The − sign indicates energy is being given off.*)

$$\lambda = \frac{c}{\nu} = \frac{c \cdot h}{\Delta E} = \frac{(3.00 \times 10^8 \cancel{\text{m}}/\text{s})(6.63 \times 10^{-34} \cancel{\text{J}} \cdot \text{s})}{1.53 \times 10^{-19} \cancel{\text{J}}} \times \frac{10^9 \text{ nm}}{1 \cancel{\text{m}}}$$

$$= 1.30 \times 10^3 \text{ nm}$$

Q. Calculate the wavelength of an electron in nanometers traveling at 61 m · s⁻¹. The mass of an electron is 9.1095 × 10⁻³¹ kg.

A. $\lambda = \frac{h}{m \cdot \upsilon} = \frac{6.63 \times 10^{-34} \cancel{\text{J}} \cdot \cancel{\text{s}}}{(9.1095 \times 10^{-31} \cancel{\text{kg}})(61 \cancel{\text{m}} \cdot \cancel{\text{s}^{-1}})} \times \frac{10^9 \text{ nm}}{1 \cancel{\text{m}}} \times \frac{(\cancel{\text{kg}} \cdot \cancel{\text{m}^2}) \cdot \cancel{\text{s}^{-2}}}{1 \cancel{\text{J}}}$

$$= 1.2 \times 10^4 \text{ nm}$$

Multiple-Choice Questions

For Questions 1–5, choose from the following choices:

A. Max Planck
B. Niels Bohr
C. Werner Heisenberg
D. Louis de Broglie
E. Wolfgang Pauli

1. No two electrons in an atom can have the same set of four quantum numbers.

2. The theory that electrons travel in discrete orbits around the atom's nucleus, with the chemical properties of the element being largely determined by the number of electrons in the outer orbits. The idea that an electron could drop from a higher-energy orbit to a lower one, emitting a photon.

3. Wave-particle duality of nature.

4. Energy can only be absorbed or released in whole-number multiples of $h \cdot \nu$.

5. The simultaneous determination of both the position and momentum of a particle has an inherent uncertainty, the product of these being not less than a known constant.

6. How many electrons can be accommodated in all the atomic orbitals that correspond to the principal quantum number 4?

A. 2
B. 8
C. 18
D. 32
E. 40

7. The four quantum numbers that describe the valence electron in the cesium atom are

A. 6, 0, –1, ½
B. 6, 1, 1, ½
C. 6, 0, 0, ½
D. 6, 1, 0, ½
E. 6, 0, 1, –½

8. Which of the following does NOT represent a possible set of quantum numbers?

A. 2, 2, 0, ½
B. 2, 1, 0, –½
C. 4, 0, 0, –½
D. 3, 2, 0, ½
E. 4, 3, 1, ½

9. A photon was found to have a frequency of 3.00×10^{14} sec^{-1}. Calculate the wavelength of the photon given that the speed of light is 3.00×10^{8} m · s^{-1} and 1 meter = 10^{9} nm.

A. 1.00×10^{-6} nm
B. 3.00×10^{-3} nm
C. 1.00×10^{3} nm
D. 3.00×10^{3} nm
E. 3.00×10^{22} nm

10. Which of the following only has 1 electron in a *p* orbital?

A. carbon
B. fluorine
C. hydrogen
D. nitrogen
E. aluminum

11. Which of the following elements most readily shows the photoelectric effect?

A. noble gases
B. alkali metals
C. halogens
D. transition metals
E. chalcogens

12. An energy value of 3.313×10^{-19} joules is needed to break a chemical bond. What is the wavelength of energy needed to break the bond?

A. 5.00×10^{18} cm
B. 1.00×10^{15} cm
C. 2.00×10^{5} cm
D. 6.00×10^{-5} cm
E. 1.20×10^{-8} cm

13. Which of the following equations represents the energy of a single electron in a hydrogen atom when it is in the *n* = 3 state?

A. $\left(\frac{-2.178 \times 10^{-18}}{3}\right)$ joules

B. $\left(\frac{2.178 \times 10^{-18}}{3}\right)$ joules

C. $\left(\frac{3}{-2.178 \times 10^{-18}}\right)$ joules

D. $\left(\frac{-2.178 \times 10^{-18}}{9}\right)$ joules

E. $\left(\frac{-2.178 \times 10^{-18}}{2^{3}}\right)$ joules

14. What is the electron configuration of tin (Sn) in the ground state in order of filling orbitals from low energy to high energy?

A. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^2$
B. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^2$
C. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^15p^3$
D. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^1$
E. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^4$

15. Electrons will enter empty orbitals of equal energy when they are available. This is (the)

- **A.** Heisenberg Uncertainty Principle
- **B.** Hund's Rule
- **C.** Pauli Exclusion Principle
- **D.** de Broglie hypothesis
- **E.** Schrödinger equation

16. The electron configuration for V^{3+} would be

- **A.** $[Ar]3d^1$
- **B.** $[Ar]3d^2$
- **C.** $[Ar]3d^3$
- **D.** $[Kr]3d^1$
- **E.** $[Kr]3d^2$

17. The electron configuration for gold in its ground state would be

- **A.** $[Au]6s^1 4f^{14} 5d^{10}$
- **B.** $[Xe]6s^2 5d^9$
- **C.** $[Xe]6s^1 4f^{14} 5d^{10}$
- **D.** $[Xe]5s^1 4d^{10}$
- **E.** $[Xe]5s^2 4d^9$

18. A Co^{3+} ion has _____ unpaired electron(s) and is _____.

- **A.** 1, diamagnetic
- **B.** 3, paramagnetic
- **C.** 3, diamagnetic
- **D.** 4, paramagnetic
- **E.** 10, paramagnetic

For Questions 19–20:

- I. carbon
- II. selenium
- III. calcium
- IV. sulfur
- V. germanium

19. Which two would have a valence electron configuration of ns^2np^4?

- **A.** I and II
- **B.** II and III
- **C.** III and IV
- **D.** II and IV
- **E.** IV and V

20. Which two would have a valence electron configuration of ns^2np^2?

- **A.** I and II
- **B.** II and III
- **C.** III and IV
- **D.** II and IV
- **E.** I and V

Free-Response Questions

1.

(a) Write the:

(i) complete ground-state electron configuration for an arsenic atom; and

(ii) use the noble gas core abbreviation method for electron configuration.

(b) Assuming arsenic to be in its ground state, write the four quantum numbers for:

(i) each of the two electrons in the 4*s* orbital; and

(ii) each of the three electrons in the 4*p* orbitals.

(c) Is the arsenic atom in its ground state diamagnetic or paramagnetic? Explain the difference.

(d) Write the formulas for and explain how the electron configuration of arsenic in its ground state is consistent with:

(i) sodium arsenide

(ii) arsenic(III) chloride

(iii) arsenic(V) fluoride

2. Answer the following questions about the element Cobalt, Co (atomic number = 27).

(a) Samples of natural cobalt show that there is 1 naturally occurring isotope, atomic mass = 59 while 33 other isotopes of cobalt have been discovered ranging from atomic mass of 48 to an atomic mass of 75. In terms of atomic structure, explain what these isotopes have in common and how they differ.

(b) Write the complete electron configuration for the cobalt atom (*e.g.*, $1s^2 2s^2$ . . . etc.) and the Co^{3+} ion.

(c) Indicate the number of unpaired electrons in the ground-state atom and the Co^{3+} ion and explain your reasoning.

(d) In terms of atomic structure, explain why the first ionization energy of cobalt (758 kJ/mol) is less than that of krypton (1351 kJ/mol) but greater than that of potassium (419 kJ/mol).

(e) Account for at least three properties of transition metals such as cobalt.

(f) Write the quantum numbers for at least three *d* electrons in the Co^{3+} ion.

Answers and Explanations

Multiple Choice

1. **E.**
2. **B.**
3. **D.**
4. **A.**
5. **C.**
6. **D.** A principal quantum number of 4 tells you that you are in the 4th energy level. The 4th energy level contains electrons in the *s*, *p*, *d* and *f* orbitals. Counting the maximum numbers of electrons available in each of the four types of sublevels; 2 in the *s*, 6 in the *p*, 10 in the *d*, and 14 in the *f* yields a total of 32. Alternatively, you can use the equation, $2n^2$ ($2 \cdot 4^2 = 32$).
7. **C.** The valence electron for the cesium atom is in the 6*s* orbital. In assigning quantum numbers, n = principal energy level = 6. The quantum number l represents the angular momentum (type of orbital). In this case, $l = 0$. The quantum number m_l is known as the magnetic quantum number and describes the orientation of the orbital in space. For *s* orbitals (as in this case), m_l always equals 0. The quantum number m_s is known as the electron spin quantum number and can take only two values, +½ and –½.
8. **A.** If $n = 2$, then l must be 0 or 1 (representing either an *s* or *p* orbital).
9. **C.** $\lambda = \frac{c}{v} = \frac{3.00 \times 10^{8}\ \cancel{\text{m}} \cdot \cancel{\text{s}^{-1}}}{3.00 \times 10^{14}\ \cancel{\text{s}^{-1}}} \times \frac{10^{9}\ \text{nm}}{1\ \cancel{\text{m}}} = 1.00 \times 10^{3}\ \text{nm}$
10. **E.** $1s^22s^22p^63s^23p^1$
11. **B.** Alkali metals have only one electron in their valence shells and thus have the lowest threshold values that are susceptible to the photoelectric effect.
12. **D.** $\lambda = \frac{h \cdot c}{E} = \frac{6.626 \times 10^{-34}\ \cancel{\text{J}} \cdot \cancel{\text{sec}} \cdot 3.00 \times 10^{10}\ \text{cm} \cdot \times \cancel{\text{sec}^{-1}}}{3.313 \times 10^{-19}\ \cancel{\text{J}}} = 6.00 \times 10^{-5}\ \text{cm}$
13. **D.** The equation is $E_n = \frac{-2.178 \times 10^{-18}}{n^2}$ joules

    where E_n = energy of the electron and n = principal quantum number
14. **A.** The electron configuration follows the Aufbau process; there are no exceptions with tin.
15. **B.**
16. **B.**
17. **C.** Stability is achieved when the outer *d* orbitals are completely filled. The 5*d* orbitals are far enough away from the nucleus where small shifts in orbital location can occur that will favor stability.
18. **D.** The electron configuration for the Co^{3+} ion is $1s^22s^22p^63s^23p^63d^6$ and would have 24 electrons; 10 pairs of electrons and four unpaired electrons in the 3*d* orbits. Atoms in which one or more electrons are unpaired are paramagnetic.
19. **D.** Sulfur: $[Ne]3s^23p^4$; Selenium: $[Ar]4s^23d^{10}4p^4$
20. **E.** Carbon: $[He]2s^22p^2$; Germanium: $[Ar]4s^23d^{10}4p^2$

Free Response

Question 1

Maximum Points for Question 1
Part (a): 2 points
Part (b): 2 points
Part (c): 2 points
Part (d): 6 points
Total points: 12

1. *(a) Write the:* *(i) complete ground-state electron configuration for an arsenic atom:* $1s^2\ 2s^2 2p^6\ 3s^2 3p^6\ 4s^2 3d^{10} 4p^3$ *(ii) use the noble gas core abbreviation method for electron configuration:* $[Ar]4s^2 3d^{10} 4p^3$	***2 points possible*** *1 point for correctly writing the complete electron configuration.* *1 point for correctly writing the noble gas core abbreviation method for the electron configuration for arsenic.*
(b) Assuming arsenic to be in its ground state, write the four quantum numbers for: *(i) each of the two electrons in the 4s orbital:* 4*s* 4, 0, 0, $+\frac{1}{2}$ and 4, 0, 0, $-\frac{1}{2}$ *(ii) each of the three electrons in the 4p orbitals:* 4*p* electrons 4, 1, −1, $+\frac{1}{2}$; 4, 1, 0, $+\frac{1}{2}$ and 4, 1, +1, $+\frac{1}{2}$	***2 points possible*** *1 point for correctly writing the electron configuration for an electron in the 4s.* *1 point for correctly writing the electron configuration for an electron in a 4p orbital.*
(c) Is the arsenic atom in its ground state diamagnetic or paramagnetic? Explain the difference. Paramagnetic. It has three unpaired electrons.	***2 points possible*** *1 point for correctly identifying that the arsenic atom is paramagnetic.* *1 point for explaining why it is paramagnetic.*

(d) Write the formulas for and explain how the electron configuration of arsenic in its ground state is consistent with: *(i) sodium arsenide* Na_3As: each Na gives up one electron to the As; the As has a complete octet; and the sodium atoms are ionically bonded to the arsenic. *(ii) arsenic(III) chloride* $AsCl_3$: the three chlorines each have one half-filled orbital and the arsenic has three. Three covalent bonds are created and the As has one non-bonding pair which forms a pyramidal structure. *(iii) arsenic(IV) fluoride* AsF_5: fluorine is highly electronegative. As such, it draws the two electrons of the non-bonding pair of $AsCl_3$ into bonding. A 4*d* orbital is involved in the sp^3d hybridization, and results in a trigonal bipyramidal geometry.	***6 points possible*** *1 point for each formula written correctly (3 points maximum).* *1 point each for a correct explanation of the bonding characteristics in each of the three examples (3 points maximum).*

Question 2

Maximum Points for Question 2
Part (a): 1 point
Part (b): 2 points
Part (c): 2 points
Part (d): 1 point
Part (e): 3 points
Part (f): 1 point
Total points: 10

2. Answer the following questions about the element Cobalt, Co (atomic number = 27). *(a) Samples of natural cobalt show that there is 1 naturally occurring isotope, atomic mass = 59 while 33 other isotopes of cobalt have been discovered ranging from atomic mass of 48 to an atomic mass of 75. In terms of atomic structure, explain what these isotopes have in common and how they differ.* The isotopes have the same number of protons (27) but different numbers of neutrons.	***1 point possible*** *No comment about the number of electrons is necessary.*

(b) Write the complete electron configuration for the cobalt atom (e.g., $1s^22s^2$ . . . etc.) and the Co^{3+} ion. Co: $1s^22s^22p^63s^23p^64s^23d^7$ Co^{3+}: $1s^22s^22p^63s^23p^63d^6$	***2 points possible*** *1 point for Co.* *1 point for Co^{3+}.*
(c) Indicate the number of unpaired electrons in the ground-state atom and the Co^{3+} ion and explain your reasoning. Co: 3 unpaired 4*s* (↑↓) 3*d* (↑↓) (↑↓) (↑) (↑) (↑) Co^{3+}: 4 unpaired 4*s* (↑↓) 3*d* (↑↓)(↑)(↑)(↑)(↑)	***2 points possible*** *1 point for Co.* *1 point for Co^{3+}.*
(d) In terms of atomic structure, explain why the first ionization energy of cobalt (758 kJ/mol) is less than that of krypton (1351 kJ/mol) but greater than that of potassium (419 kJ/mol). Ionization energy is the measure of how difficult it is to remove an electron from an atom. Since energy must be absorbed to bring about ionization, ionization energies are always positive quantities. The first ionization energy is the energy change required for the removal of the outermost electron from a gaseous atom to form a +1 ion. In general, the larger the ionization energy, the more difficult it is to remove an electron. Ionization energies increase from left to right across a period as more protons are being added (greater attraction of the electrons to the nucleus) and decrease down a group (the outermost electrons are moving farther away from the nucleus). While both potassium and krypton are in the same period as cobalt, potassium has 19 protons and krypton has 36 compared to cobalt with 27.	***1 point possible*** *1 point possible for properly explaining what ionization energy is and the reasons why ionization energy increases from left to right within a period.*
(e) Account for at least three properties of transition metals such as cobalt. Moving from left to right across the periodic table, the five *d* orbitals for transition metals become more filled. The *d* electrons are loosely bound, which contributes to the high electrical conductivity and malleability of the transition elements. The transition elements have low ionization energies. They exhibit a wide range of oxidation states or positively charged forms. The positive oxidation states allow transition elements to form many different ionic and partially ionic compounds. The formation of complexes causes the *d* orbitals to split into two energy sublevels, which enables many of the complexes to absorb specific frequencies of light. Thus, the complexes form characteristic colored solutions and compounds. Complexation reactions sometimes enhance the relatively low solubility of some compounds.	***3 points possible*** *1 point for each property.*

(f) Write the quantum numbers for at least three d electrons in the Co^{3+} ion. 3,2,2,½; 3,2,2,–½; 3,2,1,½; 3,2,0,½; 3,2,–1,½; 3,2,–2,½	***1 point possible*** *Given any three, all three must be correct.* *Other combinations are possible.*