Name: ______________________

PERIODIC PROPERTIES

1. Using the periodic table shown below indicate general electron configuration and the number of electrons in the valence shells for elements in each of the shaded groups.

Periodic Table of the Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII	VIII	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1 **H** 1.008																	2 **He** 4.00
2	3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.88	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
6	55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.8	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** 226.0	89 **Ac** 227.0	104 (261)	105 (262)	106 (263)												

Lanthanides	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.2	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0
Actinides	90 **Th** 232.0	91 **Pa** 231.0	92 **U** 238.0	93 **Np** 237.0	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

Group IA has the general electron configuration of ns^1. All elements in Group IA have one valence electron.

Group IIA has the general electron configuration of ns^2. All elements in Group IA have two valence electron.

Group IIIA has the general electron configuration of ns^2np^1. All elements in Group IA have three valence electron.

Group IVA has the general electron configuration of ns^2np^2. All elements in Group IA have four valence electron.

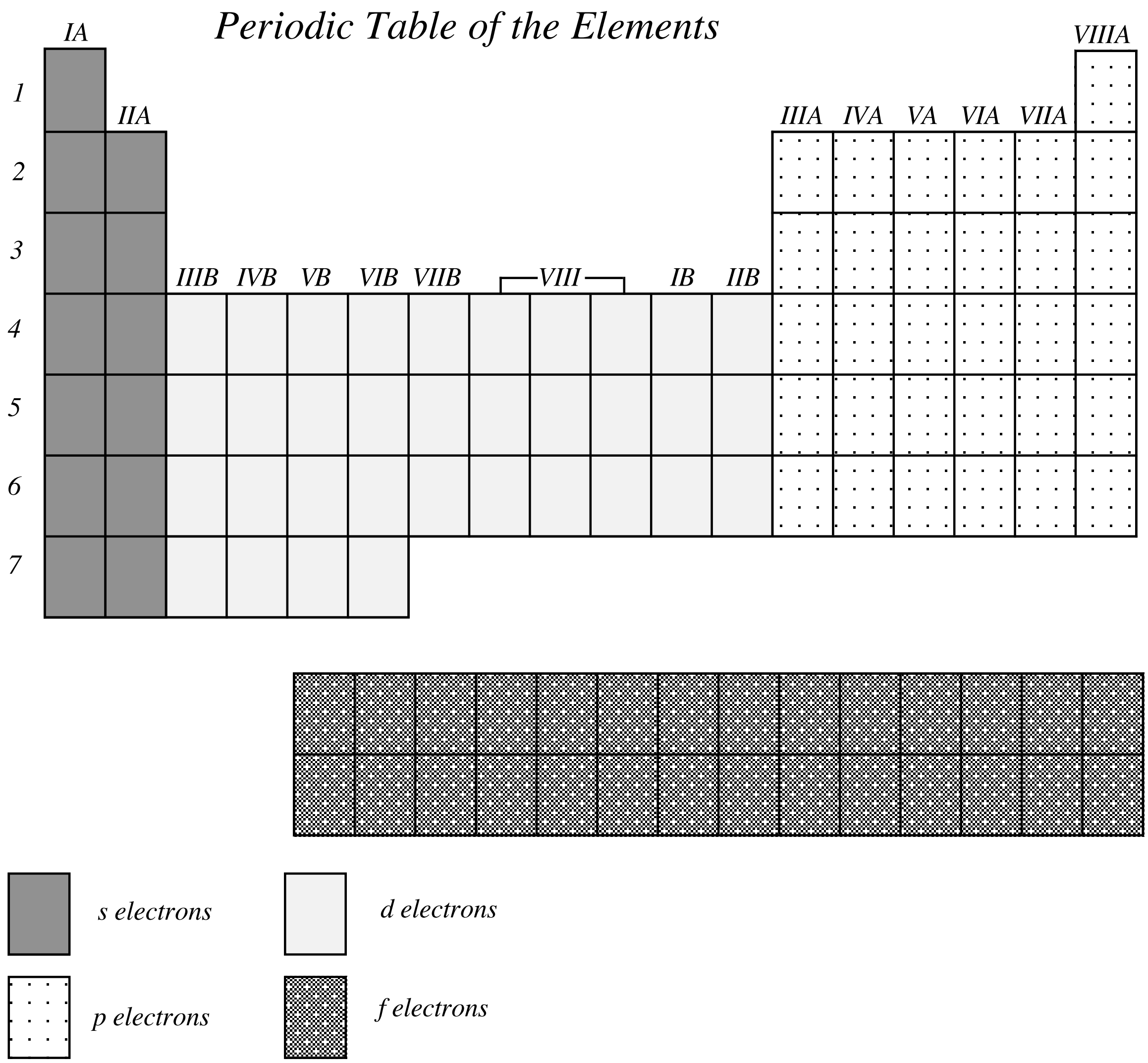
Group VA has the general electron configuration of ns^2np^3. All elements in Group IA have five valence electron.

Group VIA has the general electron configuration of ns^2np^4. All elements in Group IA have six valence electron.

Group VIIA has the general electron configuration of ns^2np^5. All elements in Group IA have seven valence electron.

Group VIIIA has the general electron configuration of ns^2np^6. All elements in Group IA have eight valence electron.

2. Use the blank periodic table below to mark the sections of the periodic table that contain elements having valence electrons in *s* orbitals, *p* orbitals, *d* orbitals and *f* orbitals.



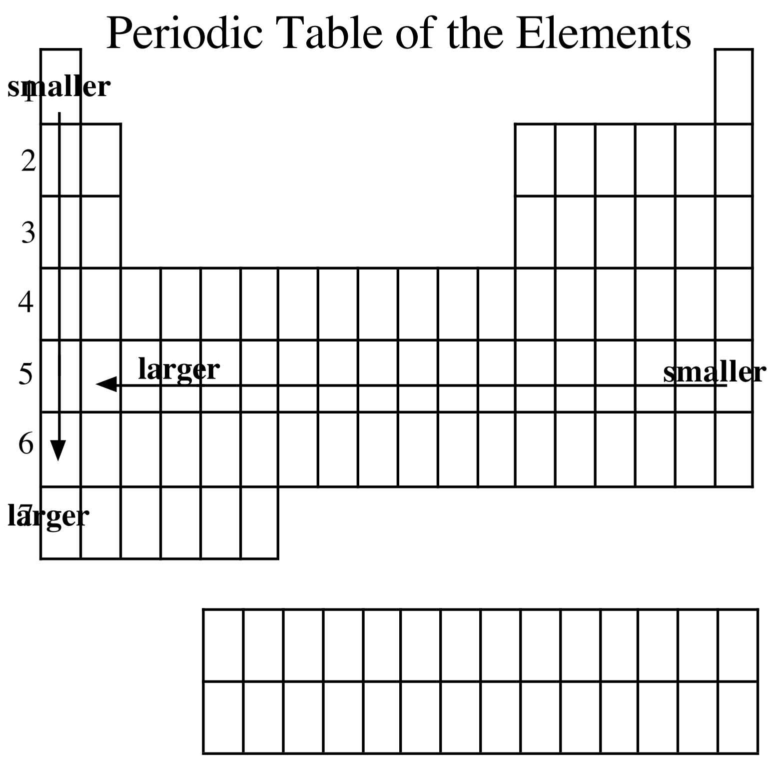
3a. Correlate group and period designations in the periodic table to the sections of the periodic table having valence electrons in *s* orbitals, *p* orbitals, *d* orbitals and *f* orbitals.

Groups IA and IIA contain elements in which electrons are filling *s* orbitals. Groups IIIA through VIIIA contain elements in which electrons are filling *p* orbitals. Groups IB through VIIIB contain elements in which electrons are filling *d* orbitals. The lanthanides and the actinides contain elements in which electrons are filling *f* orbitals.

b) Write the electron configuration for Pm and for Bi.

Pm: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 5d^1 4f^4$ **or [Xe]** $6s^2 5d^1 4f^4$
Bi: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^3$ **or [Xe]**$6s^2 4f^{14} 5d^{10} 6p^3$

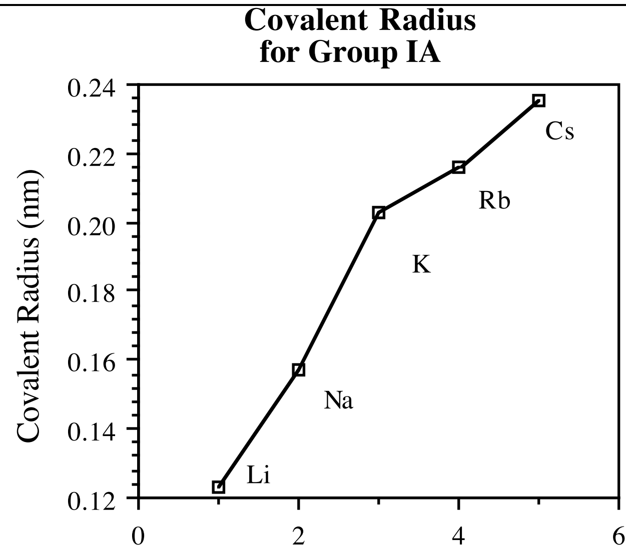
4a. Use the blank periodic table below to display the general trends in atomic radii for elements within both groups and periods (rows). Explain the observed trends in terms of electronic screening effects and the populations of the various energy levels.



In general the atomic radius of elements in a group increases going down the group. Going down a group, the valence electrons occupy orbitals of increasing larger principal quantum numbers. As the magnitude of the principal quantum number increases, the size of the orbital increases.

The atomic radius of elements in a row (period) decrease going from left to right. Going across a row, the core electrons remain constant and electrons are added to the valence shell. Since the nuclear charge increases going across a row and valence electrons do not effectively screen each other, the effective nuclear charge experienced by the valence electrons also increases. As a result the valence electrons experience a greater attraction to the nucleus and are drawn closer to the nucleus.

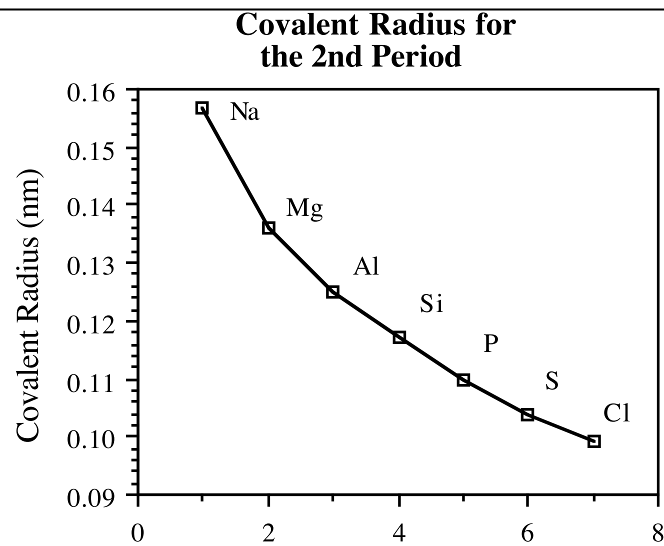
b)



Explain the trend in observed covalent radii shown in the adjacent plot.

The radius of the atoms increases as the the atomic number of the element increases. This is due to the fact the the valence electrons have higher energy (higher n values) and are thus found farther away from the nucleus.

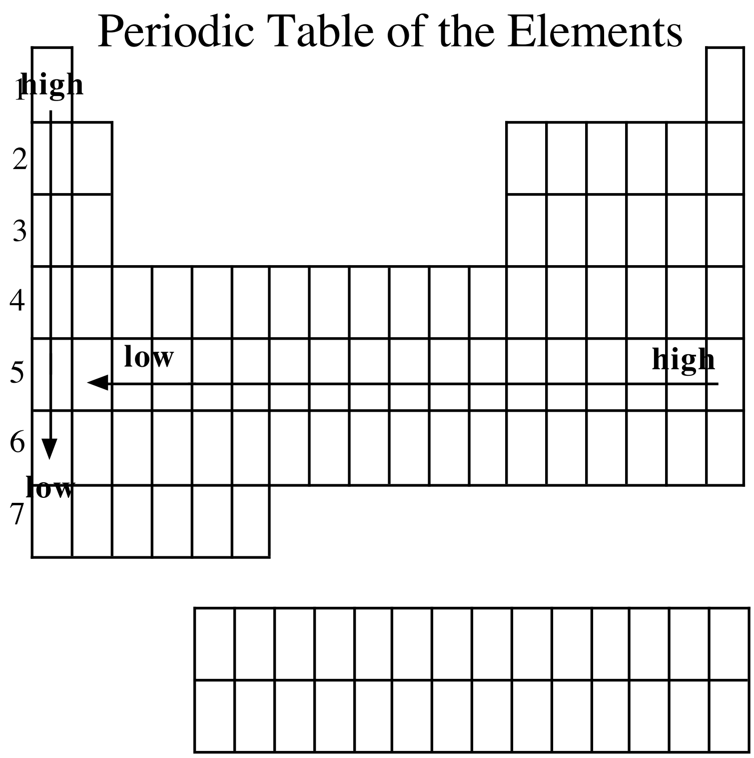
c)



Explain the trend in observed covalent radii shown in the adjacent plot.

As one moves across the 2nd period of the, energy of the valence electrons remains approximately constant, (n = 2). The nuclear charge, however, increases. The electrons are being pulled toward the nucleus more strongly on the right side of the table and thus the radii decrease as one moves from left to right.

5a. Use the blank periodic table below to display the general trends in ionization energy for elements within both groups and periods (rows). Explain the observed trends in terms of electronic screening effects and the populations of the various energy levels.

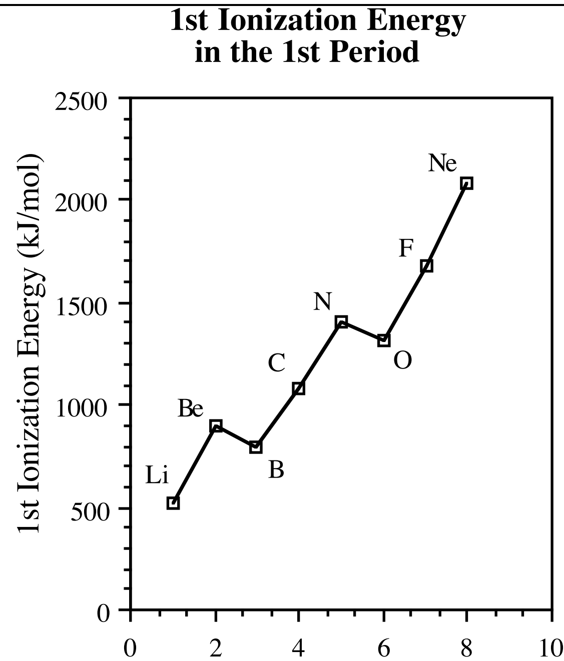


Ionization energy is the energy required to remove a single valence electron from a neutral atom.

$$M(g) \rightarrow M^{+}(g) + 1e^{-}$$

The ionization energy decreases going down a group and increases going across a period. The ionization energy decreases going down a group because the outer most electron occupies a high energy level. The higher the energy level the less energy required to remove the electron. Going across a period the electron feels a greater attraction to the nucleus and the energy required to remove the electron increases.

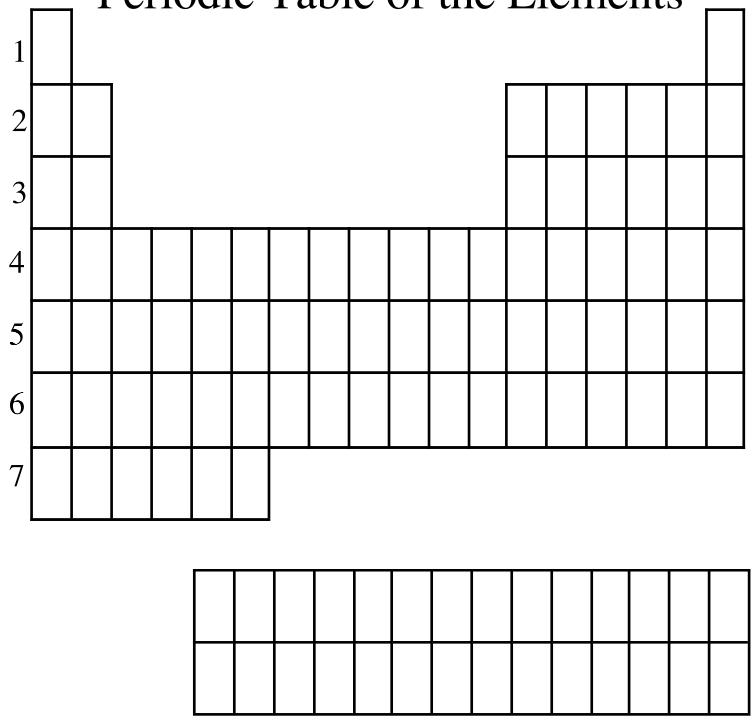
b.



In general, ionization energy increases as one moves across a period from left to right. An exception to this trend is when electrons are being removed from different subshells. The electron removed from Be comes from the 2s orbital. The electron removed from B comes from the 2p orbital. The 2p orbital is slightly higher in energy than the 2s, so the electron is easier to remove. Another exception is O. The ionization energy for O is lower than expected. This is due to the repulsion the paired electrons experience in the filled orbital. The electron which is added going to oxygen is at a higher energy level due to the repulsion is feels from the other electron and subsequently it is easier to remove.

6. Use the blank periodic table below to display the general trends in electron affinity for elements within both groups and periods (rows). Explain the observed trends in terms of electronic screening effects and the populations of the various energy levels.

Periodic Table of the Elements



Electron affinity is the energy associated with a neutral atom in the gas phase gaining an electron to form a negatively charged ion. The energy change that occurs in the following process is called the electron affinity,

$$M(g) + 1e^{-} \rightarrow M^{-}(g)$$

Electron affinity does not change significantly going down a group. The electron added is placed in the outer orbital, far away from the nucleus and it feels very little attraction to the nucleus. In addition the electron experiences a repulsion to those valence electrons already present. In general the electron affinity decreases, becoming more negative, going across a period. In general, going across a row, the valence electrons feel a greater effective nuclear charge. Addition of an electron means the new electron will feel greater attraction to the nucleus as we proceed from left to right.
There is an interesting change going from an element in Group VA to an element in Group VIA. Instead of the electron affinity decreasing it increases when moving from VA to VIA. The reason is due to the repulsion as a result of the pairing of an electron in the atom. So instead of continuing the decreasing trend at Group VIA there is a slight upswing from which it continues the downward trend.

7. Describe the observed trends in reactivity, atomic radii, ionization energy, and electron affinity within the alkali metal group.
8.

In general the reactivity of the elements in the alkali metals increases going down the group. Going down the group the atomic radius increases and the single s electron is further from the nucleus. Going down the group the valence electron is easier to remove, as indicated by the decrease in the first ionization energy. The less energy required to ionize the element the more reactive the element will be with other substances.

8. Compare atomic radii, ionization energy, and electron affinity between corresponding members of the alkali metal group (IA) and the Group IB metals.

Elements in Group IA and IB exhibit some similarity in properties. Elements in Group IA have one valence electron in an s orbital. The core electrons represent a noble gas or completely filled shell, so the loss of the single valence electron is not difficult. For the elements in Group IB, the valence electron can also be characterized as in the s orbital. However, in Group IB the increased nuclear charge means the valence electrons experience a greater attraction to the nucleus. Also, if one of the valence electrons is removed, another remains and the noble gas core is not obtained. Therefore the atomic radius is smaller in Group IB compared to Group IA and the energy required to remove an electron is greater.

9. Explain the term *degenerate*, as it relates to orbitals, and explain how adding electrons to a set of degenerate orbitals differs from adding electrons to orbitals which are not degenerate.

Degenerate orbitals are equal in energy. When adding electrons to nondegenerate orbitals, the electrons always go into the orbital of lowest energy and fill it completely. If the orbitals are degenerate, the electrons are added in accordance with Hund's rule. One electron is added to each orbital of the degenerate set before a second electron can be added to any of the orbitals.

10. Briefly explain each of the following statements.

a) The atomic radius of Mg is smaller than Ba.

Both elements are members of group IIA. The valence electrons of Ba are in the 6s orbital and those of Mg are in the *3s* orbital. Electrons in the $n = 6$ level have more energy than those in the $n = 2$ level. The orbitals they occupy are larger, so the radius of the Ba atom is greater than that of the Mg atom.

b) The ionization energy of S is lower than P.

Ionization energies tend to increase moving from left to right across rows of the periodic table. Based on this trend alone, we would expect the ionization energy of S to be greater than that of P. However, for these elements another factor must be considered. To remove an electron from P, we must disrupt a half filled set of *p* orbitals. This requires more energy than removing an electron from a subshell that is not filled or half filled. Thus, P has a higher ionization energy than S.

c) The energy required to remove the fourth electron in Al is significantly larger than the third electron.

Aluminum has an electron configuration: [Ne]$3s^2 3p^1$. There are three valence electrons that can be removed to leave the noble gas core electrons. The fourth electron would have to be removed from this noble gas core. A great deal of additional energy would have to be expended to overcome the special stability of the core electrons.

d) The energy change associated with gaining an electron, the electron affinity, is positive for some atoms and negative for others.

Electron affinity is the energy change that occurs when an electron is added to a gaseous atom. If energy is given up by the atom, then the sign of the electron affinity in negative. (The anion is more stable than the neutral atom from which it was formed.) If work must be done to force the atom to accept an electron, then the electron affinity is positive. (The anion is less stable than the neutral atom from which it was formed.)

11. Predict the products of the following reactions.

a) A sample of potassium is added to water at 25° C →

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

b) $Mg(s) + S(s) \rightarrow$

$$Mg(s) + S(s) \rightarrow MgS(s)$$

c) Sodium solid is dropped into a sample of bromine →

$$2Na(s) + Br_2(l) \rightarrow \square\square NaBr s)$$

d) $Ba(s) + O_2(g) \rightarrow$

$$2Ba(s) + O_2(g) \rightarrow 2BaO(s)$$